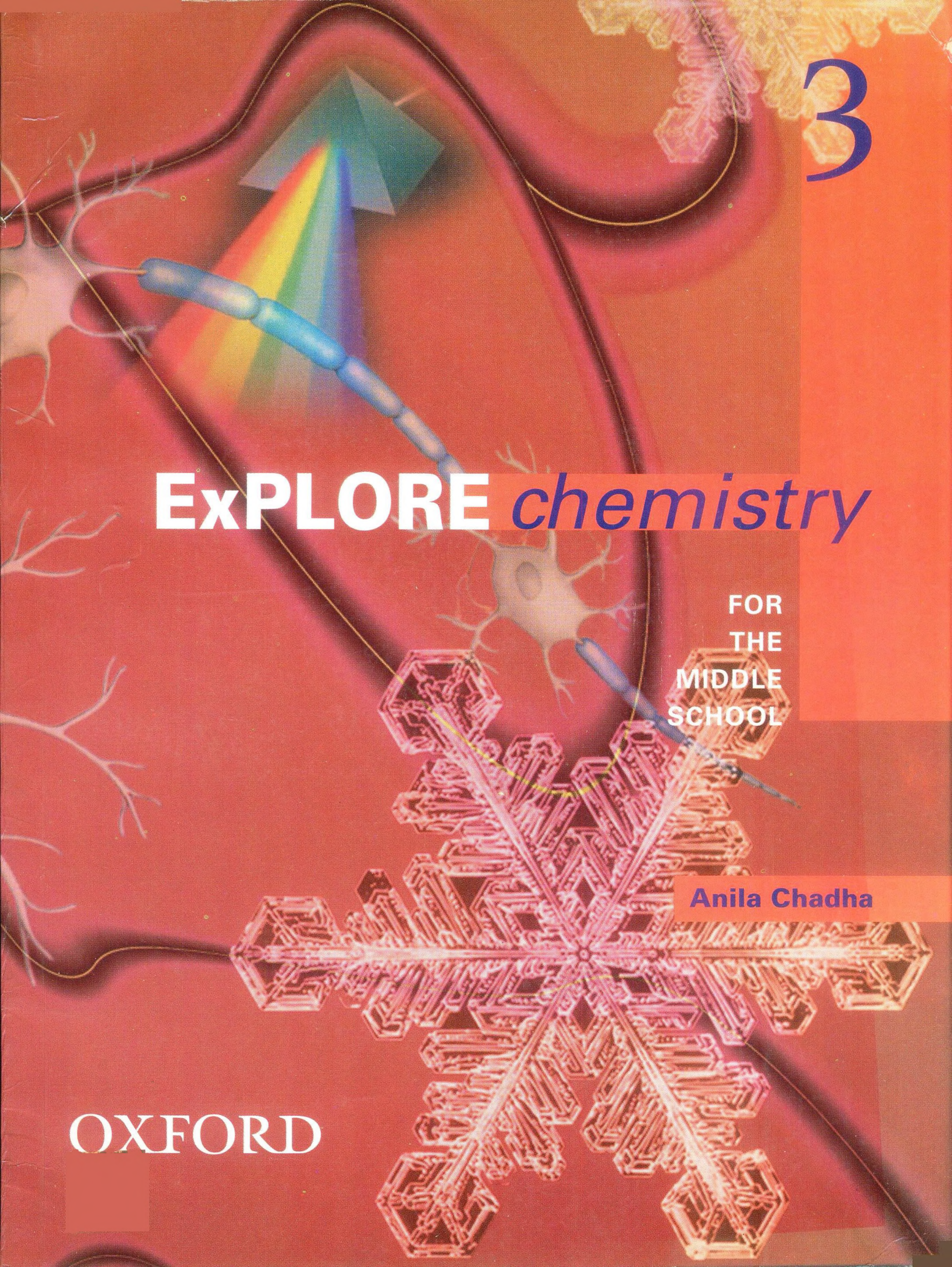

3

ExPLORE *chemistry*

FOR
THE
MIDDLE
SCHOOL

Anila Chadha

OXFORD

ExPLORE chemistry

FOR THE MIDDLE SCHOOL

3

Anila Chadha
Formerly Blue Bells School,
New Delhi

OXFORD
UNIVERSITY PRESS

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YMCA Library Building, Jai Singh Road, New Delhi 110001

Oxford University Press is a department of the University of Oxford.
It furthers the University's objective of excellence in research, scholarship
and education by publishing worldwide in

Oxford New York

Athens Auckland Bangkok Bogotá Buenos Aires Calcutta
Cape Town Chennai Dar es Salaam Delhi Florence Hong Kong Istanbul
Karachi Kuala Lumpur Madrid Melbourne Mexico City Mumbai
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Published in India
by Oxford University Press

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First published 1997
Sixth impression 2000

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ISBN 0-19-564353-4

Typeset by Anvi Composers
Printed in India at Quality Enterprises, Delhi 110006
and published by Manzar Khan, Oxford University Press
YMCA Library Building, Jai Singh Road, New Delhi 110001

Preface

Chemistry is an integral part of our lives. The clothes that we wear, the food that we eat and the air that we breathe are all governed by Chemistry.

My aim while writing this book was to make the concepts of Chemistry clear and interesting to the students. The book also aims at developing a scientific mind, logical thinking and an investigative attitude in the young learners.

The book provides a direct approach to scientific concepts and covers exhaustively the middle school syllabus. The contents have been carefully organized according to the mental level of the student, enabling him to grasp simpler concepts, facts and skills before going on to more complex ones.

The text is written in simple language. Extensive diagrams, explicit tables and schematic representations are used to illustrate the text. The book encourages the students to learn Chemistry by experimenting for themselves. Plenty of interesting activities provided in each chapter serve to instil an investigative spirit in the students and help in the acquisition of a wide variety of experimental skills.

Utmost emphasis is laid on understanding the concepts and applying them, rather than merely memorizing facts. The 'applications-to-principles' approach is adopted so that students can more readily relate chemical principles to their daily life. Recapitulation of concepts studied in previous chapters or in earlier classes, tune the students' minds to accepting the newer, more advanced ideas presented in the succeeding chapters.

Thought-provoking questions and intriguing snippets of information in every chapter serve to improve learning motivation and provide students with opportunities for active participation in the learning process.

Constructive criticism and suggestions are welcome, which would help to improve future editions of this book.

I hope that the entire series would provide interesting and stimulating learning experiences for students, and help them acquire the knowledge, skills and attitudes required for the study of Chemistry at higher levels.

Anila Chadha

Contents

1. Carbon & its allotropic forms 1
2. Compounds of carbon 13
3. Fuels, combustion & flame 25
4. Alternative sources of energy 36
5. Rocks, minerals & metals 43
6. Extraction of metals 52
7. Properties of metals 64
8. Man-made materials 77

CONTENTS—CLASS VI

Chemistry in everyday life • Chemical laboratory apparatus & equipments • Materials around us • Physical & chemical changes • Elements, compounds & mixtures • Pure substances & separation of mixtures • Water • Air around us

CONTENTS—CLASS VII

Elements, compounds & mixtures • Atoms & molecules • Chemical formulae & chemical equations • Acids, bases & salts • Water • Air & oxygen • Hydrogen • Carbon dioxide

1

carbon & its allotropic forms

- Carbon
- Allotropy
- Diamond
- Graphite
- Amorphous carbon

CARBON

The black material we see on burnt wood or toast is *carbon*. Its symbol is C. The word carbon is derived from the Latin word *carbo* meaning a black rock that can burn. Carbon is one of the most important and widely distributed elements of the earth's crust. Plants and animals are mainly made from materials containing carbon, eg, grass contains about 4 % carbon by weight. The shell of insects contain 40 % carbon.

The air around us contains tiny amounts of carbon in the form of carbon dioxide gas. Plants take in carbon dioxide through their leaves and water through their roots. Using the solar energy, they convert carbon dioxide and water into new plant material, through a process called photosynthesis. Animals like us eat plants for food, so our bodies are also partly carbon (20 %).

Carbon occurs both in the free and combined state. In the free state, it occurs as *coal*, *charcoal*, *diamond* and *graphite*. Coal contains over 90 % carbon, depending upon the type. Charcoal is almost 100 % carbon. In the combined state, carbon occurs as *carbonates* such as chalk, limestone, marble ($CaCO_3$), calamine ($ZnCO_3$), etc. When carbon combines with hydrogen, it forms a family of new materials called *hydrocarbons*. These include fuels such as natural gas, petrol, coal tar, paraffin and diesel oil. Carbon also combines with hydrogen and oxygen to form foods like sugars and starch. These are called *carbohydrates*. When foods and fuels burn, the carbon in them combines with oxygen to form carbon dioxide gas. Proteins and fats also contain carbon.

Carbon has a unique property of forming long chains of identical atoms. This is known as *catenation*. Due to this property, carbon forms a large number of compounds and a separate branch of Chemistry called *Organic Chemistry* is devoted to the study of carbon compounds. A few examples of organic compounds are starch, wax, vinegar, alcohol, perfumes, plastics, dyes, detergents, disinfectants, methane, etc.

Activity 1

Aim To show the presence of carbon in soot .

Materials required Aluminium foil, candle, match stick.

Method Place a shiny aluminium foil over a lighted candle. What happens? It gets coated with a black powder called *soot* (Fig. 1.1). What does this show? Soot contains particles of carbon in it.

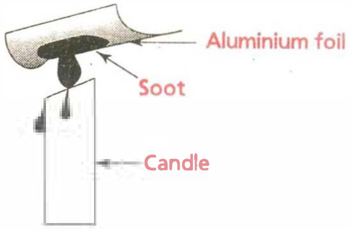


Fig. 1.1 Aluminium foil gets coated with soot

Activity 2

Aim To show the presence of carbon in sugar.

Materials required Sugar, test tube, burner.

Method Take some sugar in a test tube. Heat it over a burner. What do you notice? The sugar turns black, ie, it gets charred. This shows that sugar contains carbon particles in it (Fig. 1.2).

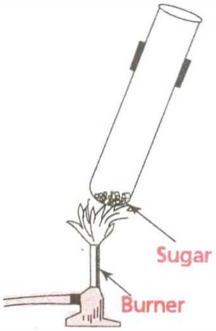


Fig. 1.2 Sugar turns black on heating

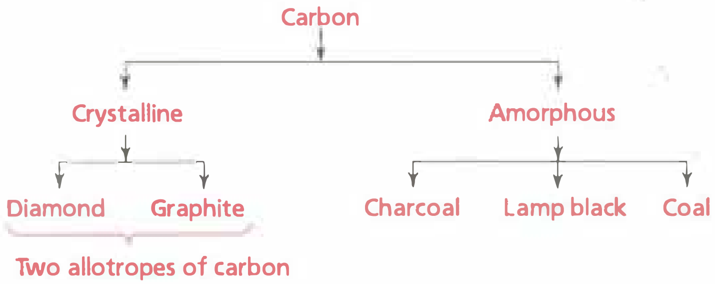
ALLOTROPY

When an element can exist in various forms which differ only in their physical properties but have the same chemical properties, the element is said to exist in its allotropic forms (allotropes) and the phenomenon is called allotropy.

Carbon, phosphorus and sulphur are some non-metals which exhibit allotropy.

The two important allotropes of carbon are *diamond* and *graphite*. Both of them are *crystalline* in nature. The other forms of carbon like *charcoal*, *coal*, *coke*, etc, contain microcrystals of graphite. They are all *amorphous* in nature.

The various forms of carbon is schematically represented below.



All these forms of carbon differ in their physical properties, but have identical chemical properties. When burnt in the presence of oxygen, all of them give carbon dioxide.

$$\text{Diamond} + \text{Oxygen} \longrightarrow \text{Carbon dioxide} + \text{Heat}$$

$$\text{Graphite} + \text{Oxygen} \longrightarrow \text{Carbon dioxide} + \text{Heat}$$

Coal + Oxygen ⟶ Carbon dioxide + Heat

Coke + Oxygen ⟶ Carbon dioxide + Heat

Remember . . .

- Carbon occurs in the free state as coal, charcoal, diamond and graphite and combined state as carbonates, hydrocarbons, proteins, carbohydrates and fats.
- Catenation is the property of carbon to form long chains of identical atoms.
- The property by virtue of which an element can exist in more than one form with similar chemical properties but different physical properties is known as allotropy.
- Diamond and graphite are the two allotropes of carbon. They are crystalline in nature.
- The amorphous forms of carbon are coal, coke, charcoal, etc.

DIAMOND

Diamond is one of the purest forms of carbon. It is represented by the symbol C.

Diamonds occur in all shapes and sizes. They were formed under intense heat and pressure, often in the core of ancient volcanoes at great depths. They are found in igneous rocks and gravels in areas as far apart as South Africa, Brazil, Australia, India, and Siberia. South Africa and Russia are the largest producers of diamonds.

Synthetic diamonds are made out of graphite, which is another pure form of carbon. Graphite is squeezed in giant presses at very high temperatures. Diamonds produced under such conditions are very minute.

Do you know ?

The weight of diamonds is measured in carats. 1 carat is equivalent to 0.2 gram. The largest diamond ever found weighed 3015 carats, or 621 g.

The diamond bearing rocks are called *Kimberlite rocks*.

In diamond, the carbon atoms are arranged in a tetrahedral manner. Each carbon atom is covalently bonded to *four* other carbon atoms. This type of bonding exists throughout the crystal to give it a three-dimensional structure. Due to the presence of strong covalent bonds, diamond is very hard. All the carbon electrons are involved in covalent bond formation. There are no free or mobile electrons, hence diamond does not conduct electricity (Fig. 1.3).

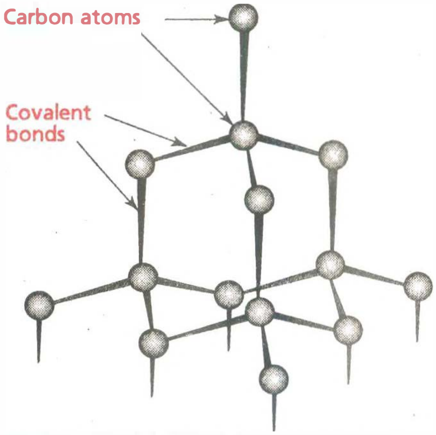


Fig. 1.3 Part of the covalent giant structure of diamond

Properties of diamond

- Diamond occurs in the free state in nature.
- It is transparent and colourless (or blue-white). Different colours are often seen in it due to the presence of traces of impurities.
- It has a high refractive index.
- It is the hardest natural substance known. One diamond can be cut only by another diamond.
- It is a non-conductor of heat and electricity.
- It has a high density.

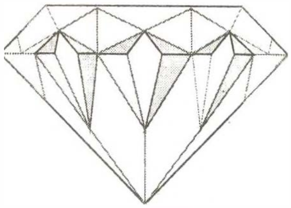


Fig. 1.4 Diamond sparkles due to the reflection of light from its cut surfaces

Uses of diamond

- Diamond sparkles brilliantly in light. So, it is used in jewellery (Fig. 1.4).
- It is used for cutting and drilling rocks.
- It is used in space missions to make radiation-proof windows for space satellites, as it can eliminate harmful radiations.

GRAPHITE

Graphite is another crystalline form of carbon. Its symbol is C. It is found in China, Sri Lanka, India, North Korea, Europe and Mexico. In India, it is found in Jammu and Kashmir, Bihar, Rajasthan, Orissa, West Bengal, Andhra Pradesh, Karnataka and Tamil Nadu.

Pure graphite can be prepared by heating coke to a high temperature of about 3000°C in an electric furnace.

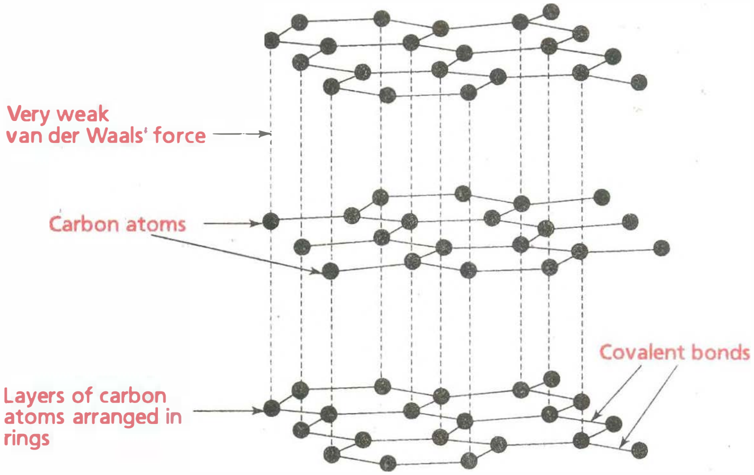


Fig. 1.5 Part of the giant structure of graphite

In graphite, the carbon atoms are arranged in the form of layers or sheets of rings of six atoms. The sheets are piled up on top of each other but there are no covalent bonds between them. They are held together only by weak forces called van der Waals' forces. The sheets can slide over one another and hence graphite is soft and greasy. Each carbon atom in a layer is covalently bonded to only *three* other carbon atoms. The fourth valence electron is free to move. This makes graphite a good conductor of heat and electricity (Fig. 1.5).

The irregular arrangement of layers does not allow light to pass through the crystal. So it is black in colour.

Properties of graphite

- It occurs in the free state in nature.
- It is soft and greasy to touch.
- It is greyish-black, opaque and a good conductor of heat and electricity.
- It is stable to heat and has a melting point of 3700 °C.

Uses of graphite

- It is used as a lubricant with petroleum jelly to form graphite grease.
- It is used as electrodes in batteries and electric furnaces.
- Carbon brushes in electric motors are made up of graphite.
- It is used for making crucibles for melting metals.
- It is used for making black core of pencils called 'pencil lead'.

For making lead pencils, graphite is crushed to a powder and then mixed either with clay or wax to form a paste. The paste is then converted to thin sticks, which on drying forms the core of our lead pencils (Fig. 1.6).

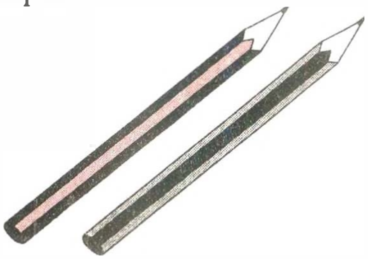


Fig. 1.6 Graphite is used in lead pencils

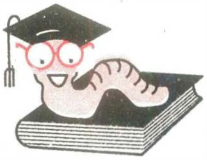
Table 1.1 Differences between diamond and graphite

	Diamond	Graphite
1.	It is transparent, colourless and has a high refractive index.	It is black, opaque and shiny.
2.	It is the hardest substance known.	It is soft and greasy to touch.
3.	It has a high density (3.5 g/cm^3).	It has a low density (2.39 g/cm^3).
4.	It is a non-conductor of electricity.	It is a good conductor of electricity.
5.	It burns in air at 900 °C to form carbon dioxide.	It burns in air at 700 °C to form carbon dioxide.

Let's Answer . . .

1. Which property allows
   (a) diamond to be used in jewellery?
   (b) graphite to be used in electrodes?
   (c) diamond to be used in cutting tools?
   (d) graphite to be used as a lubricant?

Remember . . .



- Diamond is very hard, transparent and a non-conductor of heat and electricity.
- Graphite is soft, greasy, opaque and a good conductor of heat and electricity.

AMORPHOUS CARBON

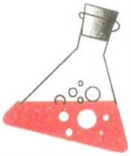
The major amorphous forms of carbon are charcoal, lamp black and coal.

Charcoal

Charcoal can be obtained from a variety of sources. It is normally named after the source from which it is obtained. It is produced by heating the source material in the absence of air. This process is known as *destructive distillation.*

Wood charcoal It is prepared by the destructive distillation of wood. It is a black, porous, brittle solid heavier than water. It floats on water as it has the capacity to hold air in its pores. It is a bad conductor of heat and electricity and a very good adsorbent.

Activity 3



Aim To prepare wood charcoal.

Materials required Test tubes, delivery tubes, trough of cold water, bunsen burner, sawdust, match sticks, stand.

Method Take some sawdust in test tube (1) fitted with a delivery tube (1). Dip one end of the delivery tube inside a test tube (2) placed in a trough of cold water. Fix another delivery tube (2) into this test tube. Heat the sawdust strongly. Bring a burning match stick near the end of the delivery tube (2). The gas liberated burns with a blue flame. It is carbon dioxide. Heat the sawdust till it chars to a black mass. What do you observe? (Fig. 1.7).

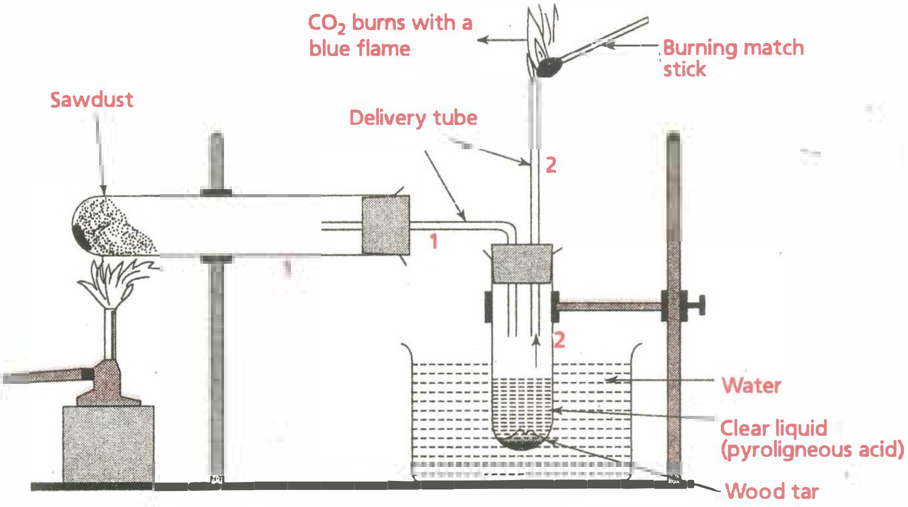

CO_2 burns with a blue flame
Burning match stick
Sawdust
Delivery tube
2
1
2
Water
Clear liquid (pyroligneous acid)
Wood tar

Fig. 1.7 Preparation of wood charcoal

A dark coloured liquid is formed at the bottom of test tube (2). It is *wood tar.* A clear liquid is formed above the dark liquid. It is known as *pyroligneous acid.* It turns blue litmus red. A black mass is formed at the bottom of test tube (1). It is *wood charcoal,* ie, the residue of sawdust.

Uses of wood charcoal

- It is used as a fuel.
- It is an important constituent of gunpowder.
- It is used in gas masks or respirators since it is porous and can adsorb harmful gases.
- It is used in making filters in cigarettes as it adsorbs the poisonous vapours of nicotine.
- It helps in decolourizing sugar syrup and refining oils and fats by removing coloured impurities.
- It is used as a disinfectant.

Animal charcoal (bone black) is obtained by the destructive distillation of bones. Bone oil and pyridine are obtained as by-products. Animal charcoal mainly contains calcium phosphate with a carbon content which is limited to 10-12 per cent only.

Animal charcoal is easily wetted by liquids and is extensively used to decolourize sugar solutions. It has the property of extracting coloured materials from solutions. It is used in the manufacture of a large number of phosphorus compounds.

Activity 4

Aim To show that animal charcoal decolourizes a sugar solution.

Materials required Two beakers, glass rod, funnel, filter paper, sugarcane juice, animal charcoal.

Method Take some sugarcane juice in a beaker. Warm it and add powdered animal charcoal to it. Heat the contents of the beaker. Now filter this solution as shown in figure 1.8. Collect the filtrate in a beaker. Do you observe any change?

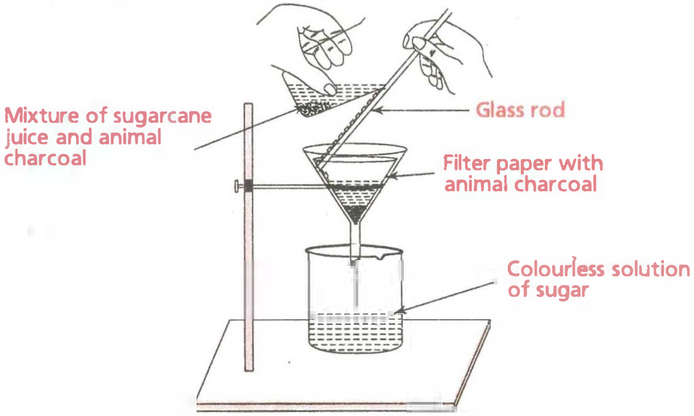

Mixture of sugarcane juice and animal charcoal
Glass rod
Filter paper with animal charcoal
Colourless solution of sugar

Fig. 1.8

The filtrate is colourless. This is because animal charcoal adsorbs the colour of the sugarcane juice on its surface and decolourizes it.

Sugar charcoal Sugar charcoal is the purest form of amorphous carbon. It is prepared by heating cane sugar or glucose in the absence of air.

$$\begin{array}{ccccc} C_6H_{12}O_6 & \xrightarrow[\Delta]{\text{Heat}} & 6C & + & 6H_2O \\ \text{(glucose)} & & \text{(sugar charcoal)} & + & \text{(water)} \end{array}$$

$$\begin{array}{ccccc} C_{12}H_{22}O_{11} & \xrightarrow[\Delta]{\text{Heat}} & 12C & + & 11H_2O \\ \text{(cane sugar)} & & \text{(sugar charcoal)} & + & \text{(water)} \end{array}$$

It can also be prepared by the dehydration of cane sugar or glucose in the presence of concentrated sulphuric acid. The acid absorbs the water to leave behind carbon.

Sugar charcoal is mostly used as a reducing agent, to derive metals from their oxides.

$$\begin{array}{ccccccc} PbO & + & C & \longrightarrow & Pb & + & CO \\ \text{(lead oxide)} & & \text{(sugar charcoal)} & & \text{(lead)} & & \text{(carbon monoxide)} \\ ZnO & + & C & \longrightarrow & Zn & + & CO \\ \text{(zinc oxide)} & & \text{(sugar charcoal)} & & \text{(zinc)} & & \text{(carbon monoxide)} \end{array}$$

Lamp black

Lamp black contains 98–99 per cent carbon.

It is prepared by heating carbon-rich substances like petroleum, kerosene, turpentine oil, etc, in a limited supply of air. The soot is collected over damp blankets in chambers and is obtained from them by jerking.

Lamp black is used as a black pigment in printing, paint and polish industries. It is also used as a filler in manufacturing rubber tyres to increase their durability.

Coal

Coal is one of the cheapest available fuels. It is a fossil fuel.

The vegetable matter buried deep under the earth millions of years ago, were decomposed by the bacteria. This decayed matter under the influence of temperature and pressure gradually changed to coal.

The main ingredients of coal are carbon, hydrogen, oxygen, nitrogen, phosphorus and sulphur. Rich deposits of coal are found in Russia, China, USA, Germany, Africa and Australia. In India, fairly large deposits are found in Jharia and Bokaro in Bihar and Raniganj in West Bengal.

Different types of coal with varying carbon contents are listed in table 1.2.

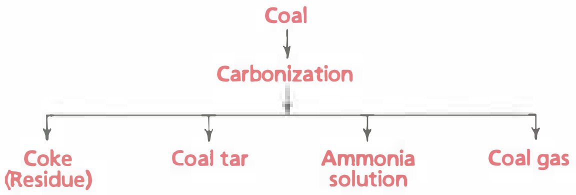
Table 1.2 Different types of coal

Types of coal	Carbon percentage (%)
1. Peat	60
2. Lignite	70
3. Bituminous	80
4. Anthracite	90

Peat has the lowest percentage of carbon, and is the inferior form of coal. *Anthracite* is the purest form of coal with the highest carbon content, and is of superior quality. *Bituminous* is the most common form of coal. It is also known as household coal.

Carbonization When coal is heated in large ovens called retorts, without letting it burn, it is called *carbonization* or *coking* (destructive distillation).

During this process, gaseous products like coal gas, ammonia solution, tar and other chemicals are released. Coke is left behind as the residue. Coke is extensively used in iron and steel production. It is also used to produce smokeless fuel and in the manufacture of electrodes.



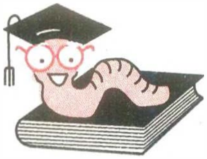
Uses of Coal

- It is used as a fuel in homes, thermal power stations, engines and furnaces, etc. It was earlier used as the source of benzene and other aromatic compounds.
- It is used to prepare coke and coal gas.
- It is used to manufacture synthetic petrol.
- The by-products of carbonization like ammonium compounds are used to produce fertilizers, dyes and explosives; tar is used in plastic and dye industry and in making adhesives.

Let's Answer . . .

1. Give reasons for the following:
   - (a) wood charcoal is used in gas masks.
   - (b) sugar charcoal is used to derive metals from their oxides.
2. Name the main product obtained by destructive distillation of the following:
   (a) wood (b) animal bones (c) coal.

Remember . . .



- Charcoal, lamp black and coal are some of the amorphous forms of carbon.
- There are three types of charcoal—wood charcoal, animal charcoal and sugar charcoal.
- Wood charcoal is prepared by the destructive distillation of wood.
- Animal charcoal is prepared by the destructive distillation of bones.
- Sugar charcoal is prepared by heating sugar or glucose in the absence of air.
- Lamp black is prepared by heating carbon-rich substances in a limited supply of air.
- Coal is a fossil fuel. Peat, anthracite, lignite and bituminous are the different forms of coal.
- Coke, coal-tar, ammonia solution and coal gas are the products of carbonization of coal.

Do you know ?

Like all materials, carbon is made of atoms. A tiny fraction of all carbon atoms are different from the rest. They are unstable. They break up and give off a burst of nuclear radiation. Such atoms are called *radioactive*.

When living things die, some of the carbon within them is trapped. Some of this trapped carbon is radioactive. Over thousands of years, the radiations from the trapped carbon get weaker and weaker. By measuring the radiation, scientists could work out the proportion of radioactive carbon. From this, they could work out how long ago death occurred. This is called *carbon dating*. Archaeologists use carbon dating to work out the age of old bones, wood, mummies, and even old cloth, because the fibres come from dead plants.

Questions

Answer in brief

1. What is allotropy? What are the different allotropic forms of carbon?
2. List the properties and uses of diamond.
3. Explain the structure of diamond and graphite.
4. Tabulate the differences between diamond and graphite.
5. Show with the help of an experiment how animal charcoal can be used to decolourize a sugar solution.
6. Give an account of the different types of coal.
7. What are the products obtained during the carbonization of coal?
8. Give the uses of lamp black.

Fill in the blanks

1. _____ is present in both living and non-living things.
2. The tendency of an element to exist in two or more forms but in the same physical state is called _____.
3. _____ chemistry deals exclusively with carbon compounds.
4. _____ and _____ are the two major allotropes of carbon.
5. _____ is the hardest substance known.

Say True or False

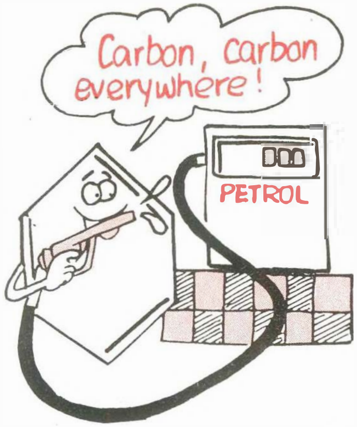
1. Charcoal is a good adsorbent.
2. Graphite has a layer structure.
3. Bone charcoal is the purest form of amorphous carbon.
4. Wood charcoal is used in gas marks and cigarette filters.
5. Coke is obtained by the destructive distillation of bones.

Choose the correct alternative

1. Anthracite is a
   (a) inferior type of coal
   (b) superior type of coal
   (c) cheapest form of coal
   (d) none of these.
2. Graphite is used
   (a) as a lubricant
   (b) in making electrodes
   (c) in lead pencils
   (d) all of these
3. Diamond is
   (a) the soft form of carbon
   (b) the hardest known substance
   (c) slippery in nature
   (d) none of these
4. Coking of coal yields
   (a) solution of ammonia
   (b) coal gas
   (c) coal tar
   (d) all of these.
5. Soot is
   (a) lamp black
   (b) charcoal
   (c) graphite
   (d) none of these.
6. Which of the following elements exhibit allotropy?
   (a) carbon
   (b) sulphur
   (c) phosphorus
   (d) all of these.
7. Charcoal is
   (a) crystalline carbon
   (b) amorphous carbon
   (c) both of these
   (d) none of these.

Think & Answer . . .

1. Why are diamonds used as precious stones for jewellery purposes?
2. Name some substances of daily use which have their origin in coal.
3. Charcoal is heavier than water, still a piece of charcoal can float in water. Why? How can a piece of charcoal be made to sink in water?
4. Firemen sometimes have to wear gas masks to protect themselves from dangerous gases. Why do these masks contain a layer of charcoal in them?



2

compounds of carbon

- Carbon monoxide
- Carbon dioxide
- Carbon cycle
- Hydrocarbons
- Methane

Carbon is present in all living and non-living matter. You have already learnt that carbon has the unique property of combining with other carbon atoms to form long chains (catenation). Due to this property, the carbon compounds in nature are unlimited.

All carbon compounds are classified under two categories, inorganic and organic compounds. Inorganic carbon compounds include the oxides, carbonates, bicarbonates, carbides, etc. Organic compounds include hydrocarbons like methane, ethane, alcohols, acids like acetic acid, methanoic acid, etc.

All organic compounds contain carbon and hydrogen. Other elements like nitrogen, oxygen, sulphur, phosphorus, magnesium, silicon may also be present in them. Let us study the preparation, properties and uses of a few carbon compounds.

CARBON MONOXIDE

Carbon monoxide was first prepared by *Lassone*, in 1776, by heating zinc oxide with charcoal. It was *Cruikshank* who established that carbon monoxide contains carbon and oxygen.

Carbon monoxide occurs in minor quantities in coal gas, tobacco fumes, volcanic gases, chimney gases, and exhaust gases of automobiles.

Preparation

Carbon monoxide is prepared by the methods mentioned below:

Reduction of carbon dioxide with carbon When carbon dioxide is passed over red hot charcoal, carbon monoxide is produced.

$$CO_2 + C \xrightarrow{\Delta} 2CO$$

Heating carbon in a limited supply of air When carbon is heated in the presence of oxygen, carbon monoxide is formed.

$$2C + O_2 \xrightarrow{\Delta} 2CO$$

Reduction of carbon dioxide with zinc or iron When carbon dioxide is passed through red hot zinc dust, carbon monoxide is produced. This method helps in the large scale production of carbon monoxide.

$$CO_2 + Zn \xrightarrow{\Delta} ZnO + CO$$

Action of steam on heated carbon or coke When steam is passed over red hot coke a mixture of carbon monoxide and hydrogen, known as *water gas*, is obtained. Carbon monoxide is separated from hydrogen by liquefaction. This method is used for the large scale production of carbon monoxide and hydrogen

$$C + H_2O \xrightarrow{\Delta} \underset{\text{(water gas)}}{H_2 + CO}$$

Physical properties

- It is a colourless and odourless gas.
- It is sparingly soluble in water and is as heavy as air.
- It condenses to a colourless liquid at –191.5 °C when cooled under pressure.
- It is extremely poisonous. If the percentage of carbon monoxide in the atmosphere increases, it results in death.

Do you know?

Haemoglobin is the substance that carries oxygen in the blood. If carbon monoxide enters the blood, it reacts with the haemoglobin in the red blood cells to form carboxy haemoglobin, which is a cherry-red poisonous compound. This inhibits the oxygen carrying capacity of blood. This is known as carbon monoxide poisoning.

A victim of such poisoning should be shifted to open air and immediately subjected to artificial respiration with carbogen—a mixture of 97% oxygen and 3% carbon dioxide.

Chemical properties

- It is neutral towards litmus solution.
- It is a non-supporter of combustion but burns in air with a blue flame to produce carbon dioxide.

$$2CO + O_2 \xrightarrow{\Delta} 2CO_2$$

- It is a good reducing agent and is used in metallurgical processes for reducing the ore to the metal.

$$PbO + CO \xrightarrow{\Delta} Pb + CO_2$$

$$ZnO + CO \xrightarrow{\Delta} Zn + CO_2$$

$$Fe_2O_3 + 3CO \xrightarrow{\Delta} 2Fe + 3CO_2$$

- It reacts with chlorine in the presence of sunlight or camphor to produce a poisonous gas, *phosgene.*

$$CO + Cl_2 \xrightarrow[\text{or camphor}]{\text{sunlight}} \underset{\text{(phosgene)}}{COCl_2}$$

- It reacts with sulphur, when passed through a hot tube, to produce *carbonyl sulphide.*

$$CO + S \xrightarrow[\text{tube}]{\text{hot}} \underset{\text{(carbonyl sulphide)}}{COS}$$

- It reacts with hydrogen when heated to 150–400 °C in the presence of zinc oxide and finely divided copper as catalyst to produce methyl alcohol.

$$CO + 2H_2 \xrightarrow{\Delta} \underset{\text{(methyl alcohol)}}{CH_3OH}$$

- It reacts with sodium hydroxide (solid) under pressure to form sodium formate.

$$CO + NaOH \xrightarrow[\Delta]{\text{pressure}} \underset{\text{(sodium formate)}}{HCOONa}$$

- It reacts with heated nickel at 50–80 °C, to form nickel carbonyl, which decomposes at a higher temperature to form nickel and carbon monoxide.

$$Ni + 4CO \underset{180°C}{\overset{50-80°C}{\rightleftharpoons}} \underset{\text{(nickel carbonyl)}}{Ni(CO)_4}$$

Test for carbon monoxide

1. It burns with a blue flame.
2. A filter paper soaked in lead (II) chloride solution turns pink, green or black due to the reduction of chloride by carbon monoxide.

Uses

- It is an important ingredient of fuels like water gas, producer gas, etc.
- It is used as a reducing agent in metallurgical processes.
- It is used in the manufacture of formic acid, methanol and synthetic petrol.
- It is used in the metallurgy of nickel.

- It is used to produce phosgene, which is used in the dye industry and in wars, and in the preparation of carbon dioxide.

Remember . . .

- Carbon monoxide is a neutral oxide of carbon.
- It is a non-supporter of combustion and an extremely poisonous gas.
- It is used as a reducing agent in metallurgy.

CARBON DIOXIDE

In 1755, Black prepared carbon dioxide by the action of dilute acids on calcium carbonate. He named it *fixed air*. Lavoisier was the first scientist to prove that carbon dioxide is an oxide of carbon.

Carbon dioxide is present in the atmosphere to the extent of 0.03 % of its volume. It ejects out from the rocks in volcanic regions and is collected in caves, mines and areas near lime kilns as *choke damp*. It occurs in mineral springs and in drinking water.

In the combined state, it is present as limestone ($CaCO_3$); magnesite ($MgCO_3$); dolomite ($CaCO_3 \, MgCO_3$); calamine ($ZnCO_3$); etc.

It is also an important product of fermentation, combustion, oxidation of organic matter, etc.

Preparation

In the laboratory, carbon dioxide is prepared by the action of dilute hydrochloric acid on marble chips or chalk or limestone.

$$\underset{\text{(calcium carbonate)}}{CaCO_3} + 2HCl \longrightarrow CaCl_2 + H_2O + \underset{\text{(carbon dioxide)}}{CO_2}$$

Other methods of preparation of carbon dioxide are discussed below.

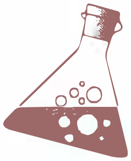
When carbon and its compounds are burnt in oxygen, carbon dioxide is obtained.

$$\underset{\text{(carbon)}}{C} + \underset{\text{(oxygen)}}{O_2} \xrightarrow{\Delta} \underset{\text{(carbon dioxide)}}{CO_2\uparrow}$$

$$\underset{\text{(methane)}}{CH_4} + \underset{\text{(oxygen)}}{2O_2} \xrightarrow{\Delta} \underset{\text{(carbon dioxide)}}{CO_2} + 2H_2O$$

A chemical reaction in which a substance burns in oxygen to produce carbon dioxide and water is known as combustion.

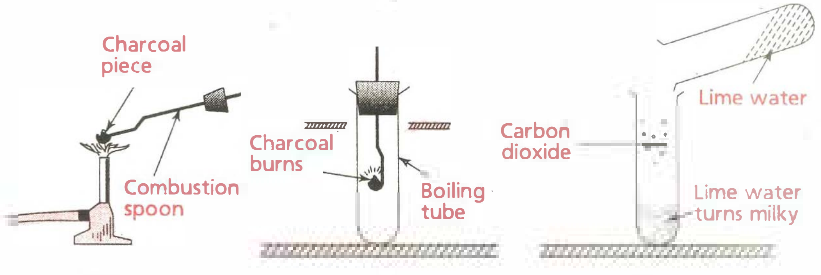
Activity 1



Aim To observe the formation of carbon dioxide gas.

Materials required Combustion spoon, boiling tube filled with oxygen, bunsen burner, piece of charcoal, test tube, lime water.

Method Place a piece of charcoal in a combustion spoon and heat it over a bunsen burner [Fig. 2.1 (a)]. The charcoal starts glowing. Now lower the combustion spoon into a boiling tube filled with oxygen [Fig. 2.1(b)]. Charcoal burns brightly and a colourless gas is given out. Remove the spoon and add lime water to the boiling tube. Lime water turns milky [Fig. 2.1 (c)]. Can you explain why?



(a) (b) (c)

Fig. 2.1

Charcoal burns in oxygen to form carbon dioxide gas. This turns lime water milky.

$$\underset{\text{(charcoal)}}{C} + O_2 \longrightarrow \underset{\text{(carbon dioxide)}}{CO_2} \xrightarrow{\text{lime water}} \text{Milky appearance}$$

By heating carbonates or bicarbonates When metal carbonates or bicarbonates are heated, carbon dioxide gas is liberated.

$$\underset{\text{(calcium carbonate)}}{CaCO_3} \xrightarrow{\Delta} \underset{\text{(calcium oxide)}}{CaO} + \underset{\text{(carbon dioxide)}}{CO_2\uparrow}$$

$$\underset{\text{(sodium bicarbonate)}}{2NaHCO_3} \xrightarrow{\Delta} \underset{\text{(sodium carbonate)}}{Na_2CO_3} + \underset{\text{(water)}}{H_2O} + \underset{\text{(carbon dioxide)}}{CO_2\uparrow}$$

By the action of dilute acids on metallic carbonates or bicarbonates Metal carbonates or bicarbonates react with dilute sulphuric acid or hydrochloric acid to form carbon dioxide.

$$\underset{\text{(sodium carbonate)}}{Na_2CO_3} + \underset{\text{(sulphuric acid)}}{H_2SO_4(\text{dil})} \xrightarrow{\Delta} \underset{\text{(sodium sulphate)}}{Na_2SO_4} + \underset{\text{(water)}}{H_2O} + \underset{\text{(carbon dioxide)}}{CO_2\uparrow}$$

$$\underset{\text{(calcium carbonate)}}{CaCO_3} + \underset{\text{(sulphuric acid)}}{H_2SO_4(\text{dil})} \xrightarrow{\Delta} \underset{\text{(calcium sulphate)}}{CaSO_4} + \underset{\text{(water)}}{H_2O} + \underset{\text{(carbon dioxide)}}{CO_2\uparrow}$$

$$\underset{\text{(sodium carbonate)}}{Na_2CO_3} + \underset{\text{(hydrochloric acid)}}{HCl(\text{dil})} \xrightarrow{\Delta} \underset{\text{(sodium chloride)}}{2NaCl} + \underset{\text{(water)}}{H_2O} + \underset{\text{(carbon dioxide)}}{CO_2\uparrow}$$

$$\underset{\text{(calcium carbonate)}}{CaCO_3} + \underset{\text{(hydrochloric acid)}}{HCl(\text{dil})} \xrightarrow{\Delta} \underset{\text{(calcium chloride)}}{CaCl_2} + \underset{\text{(water)}}{H_2O} + \underset{\text{(carbon dioxide)}}{CO_2\uparrow}$$

Physical properties

- It is a colourless and odourless gas.
- It is heavier than air (about 1.5 times denser than air).
- It is sparingly soluble in water at ordinary pressure. However, its solubility increases with increase in pressure. Hence it is used in preparing aerated drinks.
- It can be liquefied under pressure (50–60 atmosphere). Liquid carbon dioxide is stored in steel cylinders.
- When cooled below –78 °C, it directly changes into a white ice-like solid called dry ice.

Do you know ?

The commercial name for dry ice is *Drikold*. Dry ice is formed as a result of intense cooling, when carbon dioxide gas is allowed to escape from a cylinder into a woolen bag tied with a nozzle. The cooling capacity of dry ice is greater than that of ordinary ice. It forms an excellent freezing mixture, when mixed with ether or acetone (temperature as low as –110 °C). It evaporates on heating, leaving no residue. It is used as a refrigerating agent for perishable items.

Chemical properties

- Carbon dioxide is a stable gas.
- It is a non-supporter of combustion and is non-combustible. When a burning candle is introduced into a test tube containing carbon dioxide, it extinguishes (Fig. 2.2).
- It is *acidic* in nature. It dissolves in water to form carbonic acid, which turns moist blue litmus paper red.

$$\underset{\text{(carbon dioxide)}}{CO_2} + \underset{\text{(water)}}{H_2O} \underset{\longrightarrow}{\longrightarrow} \underset{\text{(carbonic acid)}}{H_2CO_3}$$

- It turns lime water milky (Fig. 2.3). This is due to the formation of insoluble carbonate. When more gas is bubbled through this solution, it becomes clear due to the formation of soluble bicarbonate.

$$\underset{\text{(lime water)}}{Ca(OH)_2} + CO_2 \longrightarrow \underset{\text{(insoluble carbonate)}}{CaCO_3} + H_2O$$

$$CaCO_3 + H_2O + CO_2 \longrightarrow \underset{\text{(soluble bicarbonate)}}{Ca(HCO_3)_2}$$

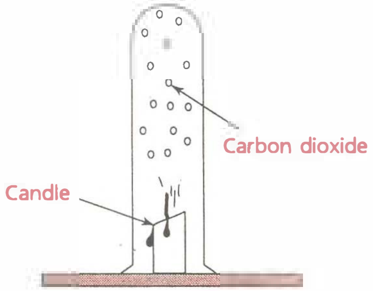


Fig. 2.2 Burning candle extinguishes in carbon dioxide

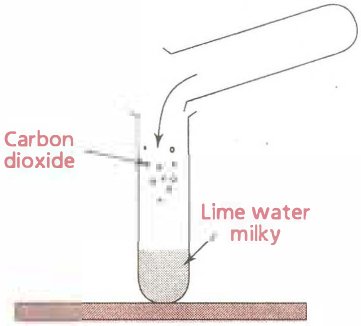


Fig. 2.3 Carbon dioxide turns lime water milky

- It reacts with sodium hydroxide solution to form sodium carbonate.

$$2NaOH + CO_2 \longrightarrow Na_2CO_3 + H_2O$$

This is an example of neutralization reaction.

Activity 2

Aim To demonstrate the neutralization reaction between sodium hydroxide and carbon dioxide.

Materials required Beaker, test tube, carbon dioxide, dilute sodium hydroxide, red litmus paper.

Method Take dilute sodium hydroxide solution in a beaker. Invert a test tube containing carbon dioxide over its surface.

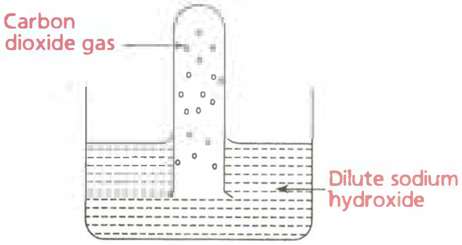


Fig. 2.4 Carbon dioxide neutralizes sodium hydroxide solution

The level of sodium hydroxide solution rises, as carbon dioxide is absorbed by it (Fig. 2.4). Dip a red litmus paper in the solution. It does not turn blue. Why?

Carbon dioxide neutralizes the sodium hydroxide solution. Now the solution is no more basic. Hence it does not turn red litmus blue.

- It is used by green plants to produce glucose, cellulose and starch in the presence of chlorophyll and sunlight. This is called *photosynthesis.*

$$6CO_2 + 6H_2O \xrightarrow[\text{(ii) sunlight}]{\text{(i) chlorophyll}} \underset{\text{(glucose)}}{C_6H_{12}O_6} + 6O_2$$

- When carbon dioxide is passed through red hot coke, it gets reduced to carbon monoxide.

$$\underset{\text{(carbon dioxide)}}{CO_2} + \underset{\text{(coke)}}{C} \longrightarrow \underset{\text{(carbon monoxide)}}{2CO}$$

Test for carbon dioxide

1. Carbon dioxide turns lime water milky.
2. It extinguishes a burning candle since it is a non-supporter of combustion.

Uses

- Carbon dioxide is used in preparing fizzy drinks.
- It is used in the manufacture of baking soda (Solvay process) and white lead.
- It is used in the manufacture of fertilizers.

- Solid carbon dioxide (dry ice) is used as a refrigerating agent. It is also used in hospitals for surgery purposes.
- Carbogen which contains 3 % carbon dioxide is used as an artificial respirator for pneumonia patients and victims of carbon monoxide poisoning.
- Carbon dioxide is used in fire extinguisher, since it is a non-supporter of combustion.

There are three types of fire extinguishers. They are:

Dry powder fire extinguisher It contains sand and baking soda ($NaHCO_3$). When this mixture is thrown over fire, baking soda decomposes to release CO_2, which extinguishes the fire.

$$2NaHCO_3 \longrightarrow Na_2CO_3 + CO_2 + H_2O.$$

Acid soda type fire extinguisher It consists of a small glass bottle filled with sulphuric acid, that is supported in a strong iron vessel containing sodium bicarbonate solution (Fig. 2.5). On striking the knob, the acid bottle breaks and carbon dioxide is liberated by the action of the acid on baking soda.

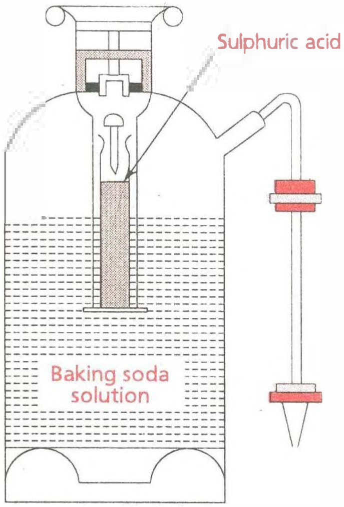


Fig. 2.5 Acid soda type fire extinguisher

Foamite type fire extinguisher In this, a solution of aluminium sulphate is taken in the glass bottle, instead of sulphuric acid. This type of fire extinguisher is specially used to extinguish oil fires.

$$Al_2(SO_4)_3 + 6NaHCO_3 \longrightarrow 2Al(OH)_3 + 3Na_2\,SO_4 + 6CO_2\uparrow$$

Let's Answer . . .

1. Which of the properties below are applicable to carbon monoxide but not to carbon dioxide?
   - (a) given out during photosynthesis.
   - (b) burns with a blue flame.
   - (c) colourless.
   - (d) poisonous.
   - (e) heavier than air.
   - (f) reducing agent in metallurgy.
   - (g) odourless.
2. Which oxide of carbon
   - (a) is used in refrigerators?
   - (b) is neutral to litmus?
   - (c) is heavier than air?
   - (d) is used to produce water gas?
   - (e) is used to produce phosgene?
   - (f) extinguishes a candle flame?
   - (g) turns lime water milky?

CARBON CYCLE

The percentage of carbon dioxide in air is around 0.03–0.05 % by volume. This value remains constant due to a series of processes taking place in the atmosphere called *carbon cycle or carbon dioxide cycle.*

Carbon dioxide is constantly produced by the combustion of fuels and respiration of plants and animals. Natural processes like decomposition of vegetable matter and fermentation also produce carbon dioxide. On the other hand, it is removed from the atmosphere during the process of photosynthesis in plants and during its dissolution in sea water or rain water. Thus, the addition of carbon dioxide in the atmosphere and its removal from the atmosphere balance each other. This balance is achieved through the carbon cycle (Fig. 2.6).

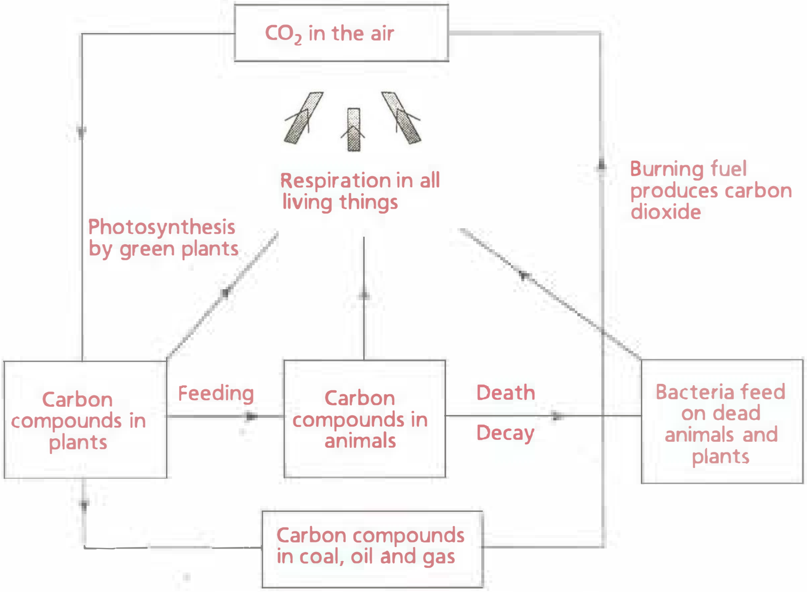

CO_2 in the air
Respiration in all living things
Burning fuel produces carbon dioxide
Photosynthesis by green plants
Carbon compounds in plants
Feeding
Carbon compounds in animals
Death
Decay
Bacteria feed on dead animals and plants
Carbon compounds in coal, oil and gas

Fig. 2.6 Carbon cycle in nature

Do you know ?

Burning fuels like petrol, oil, coal and natural gas puts about 6 billion tonnes of extra carbon dioxide into the atmosphere every year.

Remember . . .

- Carbon dioxide is an acidic oxide and a non-supporter of combustion.
- It is sparingly soluble in water and turns lime water milky.
- It is prepared by (i) complete combustion of carbon or its compounds (ii) heating carbonates and bicarbonates (iii) action of dilute acids on metal carbonates or bicarbonates.
- The amount of carbon dioxide in air remains constant. This is achieved through the carbon cycle.

HYDROCARBONS

Compounds containing carbon and hydrogen are called hydrocarbons, eg, methane, ethane, butane, petroleum, natural gas, benzene, etc.

Methane is the simplest hydrocarbon with only one carbon atom. Its formula is CH_4. Hydrocarbons containing two, three and four carbon atoms are known as *ethane, propane* and *butane*, respectively. Hydrocarbons with five or more carbon atoms are named by prefixing the Greek numerals pent (5); hex (6); hept (7), etc, to their parent names.

Table 2.1 Hydrocarbons upto ten carbon atoms

Name of the hydrocarbon	Molecular formula	Name of the hydrocarbon	Molecular formula
Methane	CH_4	Hexane	C_6H_{14}
Ethane	C_2H_6	Heptane	C_7H_{16}
Propane	C_3H_8	Octane	C_8H_{18}
Butane	C_4H_{10}	Nonane	C_9H_{20}
Pentane	C_5H_{12}	Decane	$C_{10}H_{22}$

Hydrocarbons with one to five carbon atoms are gases at room temperature, while those with more than six carbon atoms are liquids and still larger number of carbon atoms are solids at room temperature.

METHANE

Methane is the first member of the hydrocarbon series. It is also known as *marsh gas* as it is obtained from marshy places due to the decomposition of dead organic matter under water. It also exists deep in coal mines as *fire damp.* Coal gas contains 30–35 % methane. Natural gas contains 80–90 % methane.

Preparation

Methane gas is prepared in the laboratory by the action of heat on a mixture of sodium acetate and soda lime. Sodium acetate (CH_3COONa) and soda lime [NaOH mixed with $Ca(OH)_2$] are taken in a boiling tube and heated. A colourless gas is given out which is collected by the downward displacement of water (Fig. 2.7).

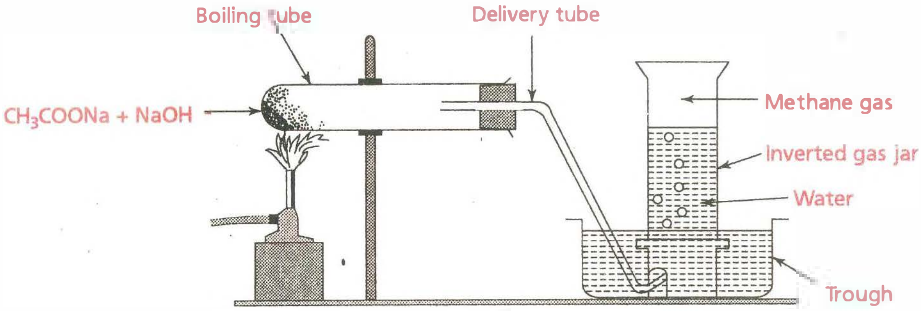


Fig. 2.7 Laboratory preparation of methane

$$\underset{\text{(sodium acetate)}}{CH_3COONa} + \underset{\text{(soda lime)}}{NaOH} \xrightarrow{\Delta} \underset{\text{(methane)}}{CH_4\uparrow} + \underset{\text{(sodium carbonate)}}{Na_2CO_3}$$

Properties

- It is a colourless, odourless and tasteless gas.
- It is lighter than air.
- It is insoluble in water but soluble in organic solvents like acetone, ether, alcohol, benzene.
- It burns in air with a bluish flame to produce carbon dioxide and water, with the evolution of heat. This reaction is known as combustion.

$$CH_4 + 2O_2 \xrightarrow{\Delta} CO_2 + 2H_2O + \text{Heat}$$

Uses

- It is used as a gaseous fuel.
- It is used in the manufacture of carbon black which finds use in printing inks, shoe polishes, gramophone records and rubber tyres.

$$CH_4 + O_2 \longrightarrow \underset{\text{(carbon black)}}{C} + 2H_2O$$

- It is used in the manufacture of methanol, chloroform, etc.

Remember . . .

- Hydrocarbons are compounds containing carbon and hydrogen.
- Methane is the simplest hydrocarbon with only one carbon atom.
- It burns in air with a bluish flame to produce CO_2, H_2O and heat.
- It is prepared in the laboratory by heating a mixture of sodium acetate and soda lime.

Questions

Answer in brief

1. Give three methods of preparation of carbon monoxide.
2. Give an account of the various methods of preparation of carbon dioxide.
3. How is methane prepared in the laboratory?
4. What are hydrocarbons? Give four examples.
5. Give two uses of carbon dioxide and carbon monoxide.
6. Write a brief note on fire extinguishers.
7. Explain carbon cycle.

8. How does carbon monoxide react with
   - (a) chlorine gas?
   - (b) sulphur?
   - (c) hydrogen?
   - (d) solid sodium hydroxide?
   - (e) red hot nickel?
9. What happens when carbon dioxide is bubbled through a solution of lime water? Write the equations involved.

Fill in the blanks

1. Carbon dioxide turns _____ milky.
2. _____ is poisonous in nature.
3. Hydrocarbons are compounds of _____ and _____.
4. _____ is called marsh gas.
5. Hydrocarbon with five carbon atoms is called _____.

Choose the correct alternative

1. Carbon dioxide is
   - (a) an acidic oxide
   - (b) mixed oxide
   - (c) neutral oxide
   - (d) basic oxide
2. Carbon monoxide is
   - (a) neutral
   - (b) alkaline
   - (c) acidic
   - (d) none of these
3. Phosgene is
   - (a) $COCl_2$
   - (b) $COCl$
   - (c) $CClO_2$
   - (d) C_2ClO
4. Carbon monoxide reacts with red hot nickel to form
   - (a) $Ni(CO)_2$
   - (b) $Ni(CO)_4$
   - (c) $Ni(CO)_3$
   - (d) $NiCO$
5. Which gas burns with a blue flame?
   - (a) CO_2
   - (b) CO
   - (c) O_2
   - (d) N_2

Think & Answer . . .

1. Carbon dioxide does not support combustion. But a burning magnesium ribbon continues to burn in CO_2. Why?
2. Respiration is the opposite reaction to photosynthesis. How?
3. Put these sentences in the correct order to make the carbon cycle complete.
   - (a) Carbon dioxide and water combine to form sugars.
   - (b) Animals give out carbon dioxide, water and a lot of energy.
   - (c) All green plants take in carbon dioxide and water.
   - (d) Animals eat plants for food.
   - (e) Carbon dioxide is present in the atmosphere.
   - (f) The sugar from the food reacts with oxygen in the cells of animals.

3

fuels, combustion & flame

- Fuel
- Combustion
- Flame

FUEL

Energy is vital to almost every human activity, from cooking to building a skyscraper. A *fuel* is a material that is used to produce heat or other forms of energy. Substances like wood, oil, natural gas, petrol, kerosene, diesel, coal, etc, are used as fuels.

All fuels burn to liberate energy. Much of the world's energy comes from the fossil fuels–coal, oil and natural gas. About 14 per cent of the world's energy is believed to come from burning wood and nearly 4 per cent from other sources. Wood is the most widely used fuel for cooking and heating.

The usefulness of a fuel is measured in terms of its *calorific value*. The amount of heat energy liberated when 1 g of the fuel is burnt completely is called the calorific value of a fuel. It is expressed in calorie/gram. Table 3.1 shows the calorific value of some common fuels.

Table 3.1 Calorific value of some fuels

Fuel	Calorific value (cal/g)
Wood	4000
Coal	7000
Coke	8000
Petrol	11500
Natural gas	8000–12000
Water gas	3000–6000
Hydrogen gas	34000

Do you know ?

- 1 g of coal contains about 25 kilo joules of energy.
- 1 g of oil contains about 45 kilo joules of energy.

Characteristics of a good fuel

A large number of substances burn to produce energy, but all these substances are not used as a fuel. The characteristics of a good fuel are:

- It should have a high calorific value.
- It should be cheap and readily available.
- It should be easy to store, transport and handle.
- It should not produce toxic fumes or smoke.
- It should not leave behind excess of ash or soot.
- It should have a low ignition temperature. [Ignition temperature is the minimum temperature at which a substance starts burning].

Types of fuels

There are three types of fuels. They are solid, liquid and gaseous fuels.

Solid fuel Coal, wood, charcoal, coke, paraffins and tallow are some common solid fuels. Wood was the first solid fuel used by man.

Solid fuels require much space for storage and leave smoke and ash on burning.

Liquid fuel Petrol, kerosene, diesel, methanol are some common liquid fuels. Most liquid fuels are obtained from petroleum. They leave no residue when burnt and can be stored easily.

Let us study about one liquid fuel in detail.

Petroleum The word petroleum is derived from the Latin word *petra* meaning rocks and *oleum* meaning oil. It is an oily mixture of hydrocarbons in its crude form, so it is called *crude oil*.

Petroleum is a versatile form of energy. In fact, the economy of a nation depends to a great extent on its petroleum wealth. Hence it is also called *black gold*.

Petroleum is a fossil fuel formed by the decomposition of plant and animal matter buried under the earth millions of years ago. It occurs deep under the earth between layers of non-porous rocks (Fig. 3.1). Oil deposits are also found under ocean beds due to the death of marine

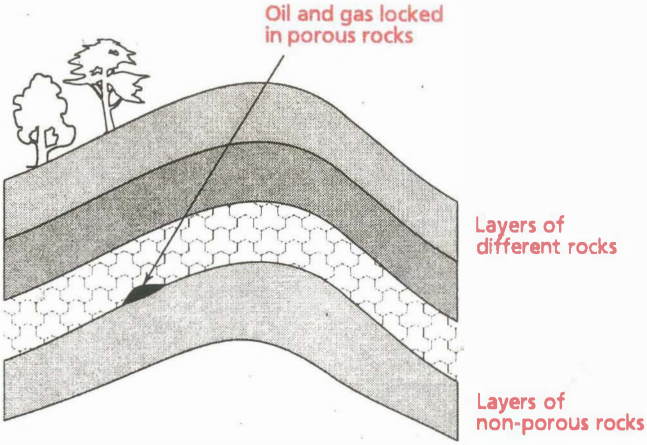

Oil and gas locked in porous rocks
Layers of different rocks
Layers of non-porous rocks

Fig. 3.1 Crude oil under the earth's surface

animals and plants like *plankton*. Natural gas occurs above the crude oil under the rocks.

The largest oil fields are located in Saudi Arabia and Kuwait. America is the largest producer of petroleum, accounting for nearly 50 % of the world's total output. In India, petroleum is produced in Ankleshwar and Kalol in Gujarat, Rudrasagar and Lakwa in Assam, Bombay High (offshore area) and deltas of Cauvery, Krishna and Godavari rivers.

Petroleum is obtained by drilling holes into the earth's crust, where the oil is found. When the oil wells are drilled, natural gas comes out first through the rocks with great pressure, after which petroleum comes out by itself. After the pressure subsides, it can be pumped out of the oil wells.

Crude oil gushing out of the wells is a dark coloured, viscous, oily liquid, made up of a mixture of hydrocarbons. It is separated by refining.

Refining is the process of separating petroleum into useful fractions by fractional distillation.

It is carried out in a tubular furnace with a long fractionating column (Fig. 3.2).

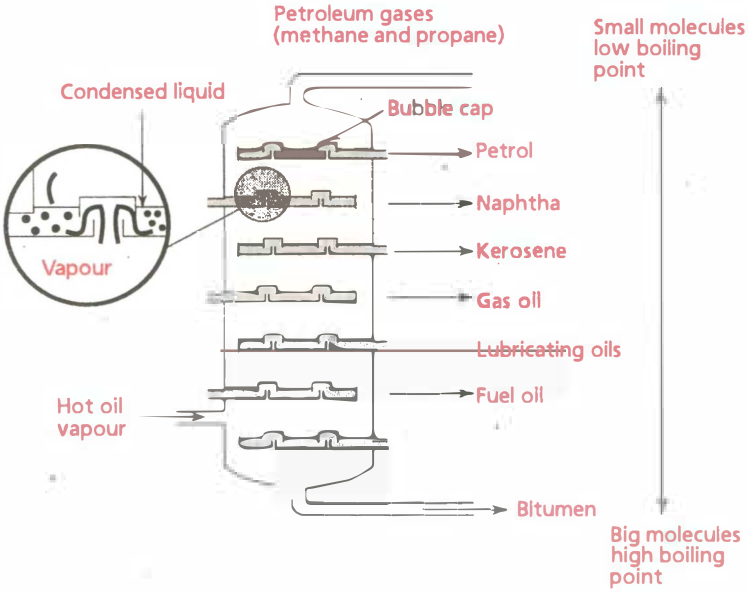


Fig. 3.2 Fractional distillation of petroleum

Crude oil is heated in a furnace to about 800 °C, and the vapours formed are passed into the fractionating column. The vapour of the higher boiling fraction condenses first while those of the lower boiling fractions rise to the top and condense later. The process is continued until the crude oil is separated into five or six fractions. Each fraction boils over a different range of temperature. Table 3.2 shows the various fractions obtained by the fractional distillation of crude oil.

Table 3.2 Various fractions obtained and their uses

Fraction	Boiling point range	Number of carbon atoms in a chain	Uses
Refinery gas dissolved in the oil	less than 40 °C	1–4	Fuels.
Petrol and naphtha	40 °C–170 °C	4–10	Fuel for cars, chemicals for plastics and insecticides.
Kerosene	150 °C–240 °C	10–16	Jet fuel, detergent.
Diesel oil	220 °C–250 °C	16–20	fuel for lorries.
Lubricating oils	250 °C–350 °C	20–30	Polishes and waxes, oil for machinery.
Fuel oil	above 350 °C	30-40	Fuel for ships and power stations.
Paraffin waxes	solid	40–50	Candles, polishes wax-papers, water proofing, grease.
Bitumen	solid	around 60	Tar for roads, weather-proofing materials.

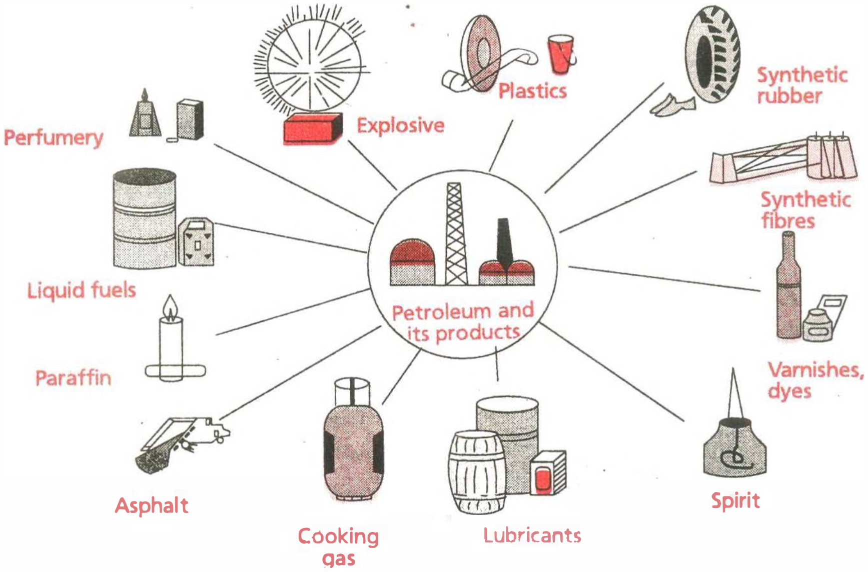


Fig. 3.3 Useful petroleum products

Gaseous fuels Natural gas, producer gas, coal gas, water gas, LPG (liquefied petroleum gas) and biogas (Gobar gas) are some examples of gaseous fuels.

Natural gas is obtained from petroleum wells and is a mixture of hydrocarbons (mainly methane). It is the cheapest available gaseous fuel.

Producer gas is a mixture of nitrogen and carbon monoxide. It is prepared by passing air over red hot coke.

$$2C + O_2 + 4N_2 \longrightarrow \underset{\text{(producer gas)}}{2CO + 4N_2}$$

Coal gas is a mixture of carbon, hydrogen and methane. It is prepared by the destructive distillation of wood.

Water gas is a mixture of hydrogen and carbon monoxide. It is prepared by passing steam over red hot coke.

$$C + \underset{\text{(steam)}}{H_2O} \longrightarrow \underset{\text{(water gas)}}{CO + H_2}$$

Liquefied petroleum gas (LPG) is a mixture of butane and isobutane, which are liquefied by compression. The domestic gas cylinders contain liquefied butane and isobutane under pressure. A strong smelling substance, ethyl mercaptan (C_2H_5SH) is added to this mixture to detect gas leakage. It is obtained by the fractional distillation of petroleum and from natural gas. It is a clean and safe fuel.

Biogas (Gobar gas) is chiefly composed of methane. It is obtained by the degradation of plant and animal matter.

Do you know

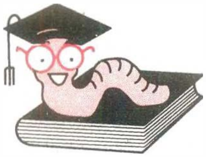
There are a number of new fuels which could be developed. One such fuel is obtained from household rubbish. This fuel contains the same amount of energy as 20 million tonnes of coal. A number of cities, including London, Paris, Amsterdam and Frankfurt, have power stations that burn household rubbish to produce electricity.

Many plants such as maize, cassava and sugarcane are rich in sugars. The sugars could be collected and converted into alcohol. Alcohol is a valuable concentrated fuel. Brazil is one country that has been trying to reduce its oil imports by growing these plants. Many cars in Brazil use alcohol instead of petrol.

Let's Answer . . .

1. Classify the following as solid, liquid and gaseous fuels:
   Petrol, coal, wood, oil, rubbish, natural gas, LPG, coke, water gas, charcoal, kerosene.
2. Coal has a calorific value of 7000 cal/g while petrol has a calorific value of 11500 cal/g. Which is a better fuel?
3. Name the oil fraction where the compound
   - (a) has the longest chain length?
   - (b) has the lowest boiling point?
   - (c) has the highest boiling point?

Remember . . .



- A fuel is a chemical which releases energy [heat or other forms of energy] when burnt.
- The amount of heat energy liberated when 1 g of the fuel is burnt completely is called the calorific value of a fuel.
- Coal, wood and charcoal are solid fuels. Petroleum, kerosene, and diesel are liquid fuels. Water gas, natural gas and LPG are gaseous fuels.
- Refining of petroleum yields gases, petrol, naphtha, kerosene, diesel oil, lubricating oils, fuel oil, bitumen and paraffin waxes.
- LPG is a mixture of butane and isobutane. It is used as household cooking gas.

COMBUSTION

Most substances when burnt produce heat and light, eg, wood, paper, kerosene, straw, wool, etc. Such substances which burn easily are called *combustible substances*. However, substances like glass, iron, stone, etc, do not burn easily. They are called *non-combustible substances.*

Burning is a chemical change. During burning, substances combine with the oxygen in air and get oxidized. Hold a piece of magnesium ribbon in a flame. It catches fire and burns with a dazzling white light.

This process of burning of a substance in the presence of oxygen with the evolution of heat and light is called combustion.

For example, charcoal burns in oxygen to produce carbon dioxide and heat.

$$C + O_2 \longrightarrow CO_2 + \text{Heat}$$

Methane burns in oxygen to produce carbon dioxide, water and heat.

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O + \text{Heat}$$

Other examples of combustion reactions are burning of a candle, a jet of hydrogen burning in chlorine, etc. They can also be called *oxidation* reactions. All fuels undergo oxidation during burning.

Can glowing of a bulb be called a combustion reaction? No, because though it produces heat and light, it does not involve any chemical change.

Combustion reactions take place under the following conditions:

- Presence of a combustible substance (substance that burns).
- Presence of a supporter of combustion (medium in which combustion takes place).
- Attainment of ignition temperature (lowest temperature at which a substance begins to burn).

All substances do not have the same ignition temperature. For example, white phosphorus, petrol, diesel have low ignition temperatures. Wood and coal have high ignition temperatures.

Types of combustion

There are three types of combustion reactions:

- Rapid combustion
- Spontaneous combustion
- Explosion.

Rapid combustion The oxidation reaction in which large amounts of heat and light are produced in a short time is called *rapid combustion*. It takes place very rapidly. For example, when a burning match stick is brought near a candle, it starts burning immediately. Other examples are burning of gas in a burner, stove, etc.

Spontaneous combustion A combustion reaction in which no external heat is taken in is called *spontaneous combustion*, eg, burning of white phosphorus in air. The ignition temperature of white phosphorus is very low (35 °C). It catches fire immediately. Hence it is always kept under water.

Activity 1

Aim To demonstrate spontaneous combustion.

Materials required Filter paper, carbon disulphide solution, white phosphorus, beaker, a pair of tongs.

Method Dissolve a few pieces of white phosphorus in carbon disulphide (CS_2) solution taken in a beaker. Dip a filter paper into the solution. Now hold the filter paper with a pair of tongs and wave it in the air (Fig. 3.4). The carbon disulphide evaporates and the filter paper catches fire. Why does this happen?

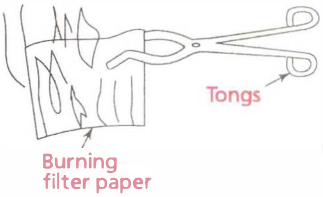


Fig. 3.4 Spontaneous combustion of white phosphorus at 35 °C

White phosphorus undergoes slow oxidation in air and heat is produced during the process. When the ignition temperature (35 °C) is reached, it starts burning in air and catches fire. Hence the filter paper bursts into flame. This is an example of spontaneous combustion.

Explosion A reaction involving release of tremendous amount of heat, light and sound in a short interval of time is called *explosion*. For example, bursting of a fire cracker when it is in contact with a burning splinter or when pressure is applied to it.

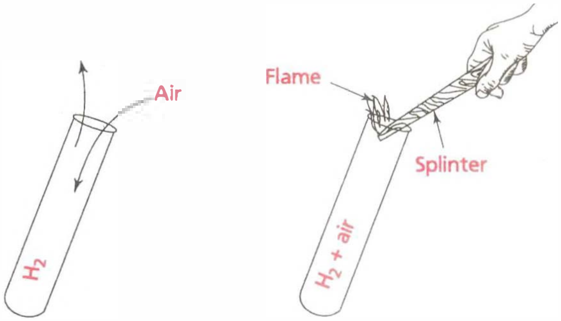


Fig. 3.5 Formation of an explosive mixture when hydrogen comes in contact with air

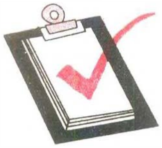
Another striking example of explosion is the combustion of hydrogen. When hydrogen is mixed with air, large amounts of heat and steam are produced which can cause explosion (Fig. 3.5). The mixture explodes with a popping sound.

$$2H_2 + O_2 \longrightarrow \underset{\text{(steam)}}{2H_2O} + \text{Heat}$$

Remember . . .

- Combustion is the process of burning a substance in the presence of oxygen to give heat and light.
- There are three types of combustion—rapid combustion, spontaneous combustion and explosion.

Let's Answer . . .



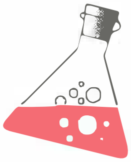
1. Arrange these substances, starting with the one that catches fire most easily.

   Wood, coal, kerosene, spirit, wax, paper.
2. If a little kerosene is poured into wood, it catches fire more easily. Why?

FLAME

A *flame* is a zone of burning gas. Some substances burn with a flame while others do not. A flame is produced when a combustible substance is in the vapour state on heating. Thus, a flame is a zone of combustion of gaseous substances with the evolution of heat and light.

Activity 2



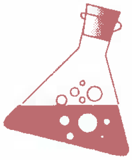
Aim To show that substances which vaporize on heating burn with a flame

Materials required Camphor, wood, charcoal.

Method Hold some pieces of camphor, charcoal and wood in a candle flame, one at a time. Camphor burns with a flame. Charcoal glows but does not produce a flame. Wood burns with a flame, but later glows like charcoal. What do you infer from this?

Camphor vaporizes on heating, but wood and charcoal do not. Only those substances which vaporize on heating burn with a flame.

Activity 3



Aim To study a candle flame.

Materials required Candle, match sticks.

Method Light the wick of a candle with a burning match stick. It burns with a flame (Fig. 3.6). How is the flame produced? When the candle is ignited, the

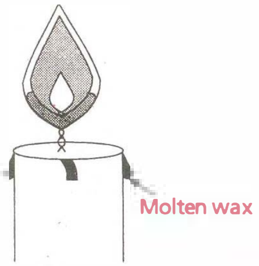


Fig. 3.6 A candle flame

wax around the wick melts. This molten wax rises up through the wick and vaporizes. The vapours burn in the oxygen of the air to produce the flame. Thus, the flame is produced by the reaction between paraffin wax (hydrocarbon) vapours and oxygen.

Parts of a flame

The different parts of a candle flame is shown in figure 3.7.

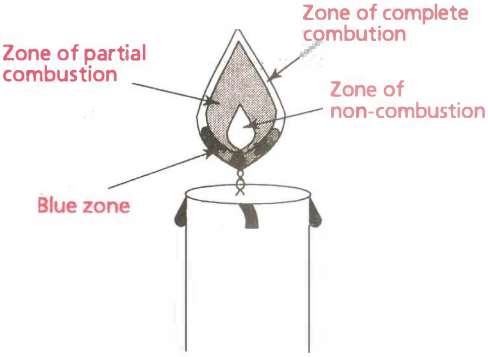


Fig. 3.7 Parts of a candle flame

Zone of non-combustion This is the dark zone that lies around the wick. It contains unburnt gas particles. No combustion takes place here as no oxygen is available.

Zone of partial combustion In this zone, the hydrocarbons present in the oil gas decompose into free carbon and hydrogen. The unburnt carbon particles impart a pale yellow colour to the flame. This is the *luminous* part of the flame.

Zone of complete combustion This is the *non-luminous* thin zone of the flame. It is the outermost hottest region in the flame that is invisible. Here, carbon and hydrogen are completely oxidized to CO_2 and water vapour.

Blue zone It lies at the bottom of the flame. The blue colour is due to the burning of CO produced due to the incomplete combustion of carbon particles.

$$C + O_2 \longrightarrow 2CO$$

(blue flame)

$$2CO + O_2 \longrightarrow 2CO_2$$

Table 3.3 Average temperatures of some common flames

Flame	Temperature (°C)	Flame	Temperature (°C)
Sun (outer part)	7000	Oxy-hydrogen blow pipe	2400
Electric arc	3500	Bunsen burner	1870

Remember . . .

- A flame is produced when a combustible substance is in the vapour state on heating.
- The various zones in a flame are
  (i) non-luminous zone (complete combustion)
  (ii) luminous zone (partial combustion)
  (iii) dark zone (no combustion)
  (iv) blue zone

Questions

Answer in brief

1. What is a fuel? What are the requirements of a good fuel?
2. How are fuels classified? Explain with examples.
3. Give the method of preparation of any two gaseous fuels.
4. How is petroleum refined? What are the various fractions obtained during the refining process? Mention their uses.
5. Under what conditions do combustion reactions take place?
6. What are the different types of combustion? Explain briefly.
7. Draw a neat labelled diagram of a candle flame and explain the various parts.

Fill in the blanks

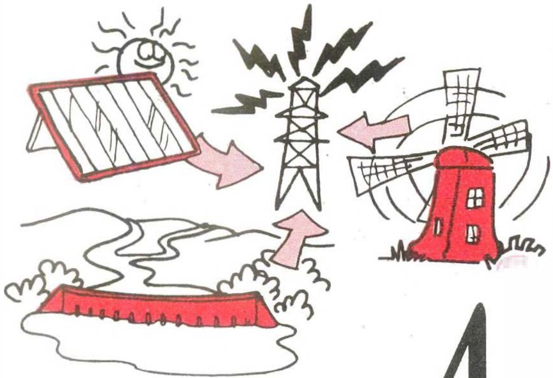
1. A good fuel should have a high _____ value and a low _____ temperature.
2. _____ and _____ are fossil fuels.
3. Ignition temperature of white phosphorus is _____.
4. Paper is a _____ substance.
5. Refining of petroleum is done by _____.
6. _____ is a chemical change that produces heat and light.
7. Burning of a candle is a _____ change.
8. Non-luminous zone is the zone of _____ combustion.

Choose the correct alternative

1. Glass is
   (a) non-combustible (b) combustible
   (c) none of these
2. Camphor
   (a) vaporizes in air (b) condenses in air
   (c) none of these
3. Carbon monoxide burns with
   (a) no flame (b) yellow flame (c) blue flame.
4. For combustion reactions
   (a) air is essential (b) combustible substance is essential
   (c) both (a) and (b)
5. Dark zone is
   (a) visible (b) invisible (c) green in colour.

Think & Answer . . .

1. When a fuel containing carbon, hydrogen and oxygen is burnt, what are the gases that will be obtained?
2. Why is cooking gas a better fuel than coal?
3. How many kilograms of coal would produce the same amount of heat as 3.5 kg of petrol on burning? (Calorific value of petrol = 11,500 cal/g; coal = 7000 cal/g)
4. Why should water not be used to put out burning petrol or oil?



4

alternative sources of energy

- Energy • Renewable energy sources
- Non-renewable energy sources
- Conservation of energy

ENERGY

Energy is the capacity to do work. Nothing can live, move or work without energy. Plants need energy to grow. We need it to keep alive. Machines need a supply of energy to work. Some machines use electricity, while others burn fuel which contains stored energy.

There are different kinds of energy. Heat, light and sound are kinds of energy which move around. Heat and light are *radiant energy*, often called radiation. When we put on an electrical switch, we use *electrical* energy. It provides lighting and heating, and makes many machines work. The energy stored in foods and fuels is *chemical* energy. It is released when fuels are burnt or food is oxidized. When atoms split or join to make different atoms, huge amounts of energy may be released. This is called *nuclear* energy. It can be used to make electricity or destructively in nuclear bombs. Apart from this, there is *mechanical* energy which includes both *kinetic* and *potential* energy. All these kinds of energy can be converted from one form to another.

Thus, energy is obtained from a variety of sources. All sources of energy can be classified under two main categories:

(i) Renewable energy sources
(ii) Non-renewable energy sources

RENEWABLE ENERGY SOURCES

Sources of energy which are produced in nature continuously and do not get exhausted with use are called renewable sources of energy. Examples–Sun, wind, water, biogas, etc.

Solar energy

Almost all our energy originally came from the sun. At the sun's centre, nuclear reactions take place. These give enormous amounts of energy which radiate from the surface of the sun as heat and light. Green plants use the solar energy to make their food from simple materials like water and carbon dioxide (photosynthesis).

Light and heat from the sun pour down on the earth all the time. When we turn this energy into electricity or use it as heat, we call it *solar power.*

Electricity from Sun Solar cells can turn sunlight directly into electricity. Solar cells are made from thin slices of pure silicon, a material that can be extracted from sand. The top of the slice is made up of a different kind of silicon, compared to the bottom. When light shines on it, an electric current flows along a wire connecting the top to the bottom. A single solar cell produces tiny current, but an array of cells connected together produce a useful amount of power.

Solar cells are expensive to make, so they are used only where there is no convenient electrical supply.

Utilizing Sun's heat The sun's energy can be trapped as heat. *Solar panels* on the roofs of buildings can trap the sun's heat and are used to heat water. In a solar panel, pipes are fixed to a black plate, under a sheet of glass. The sun heats up a liquid in the pipe, and this liquid heats up a tank of water (Fig. 4.1).

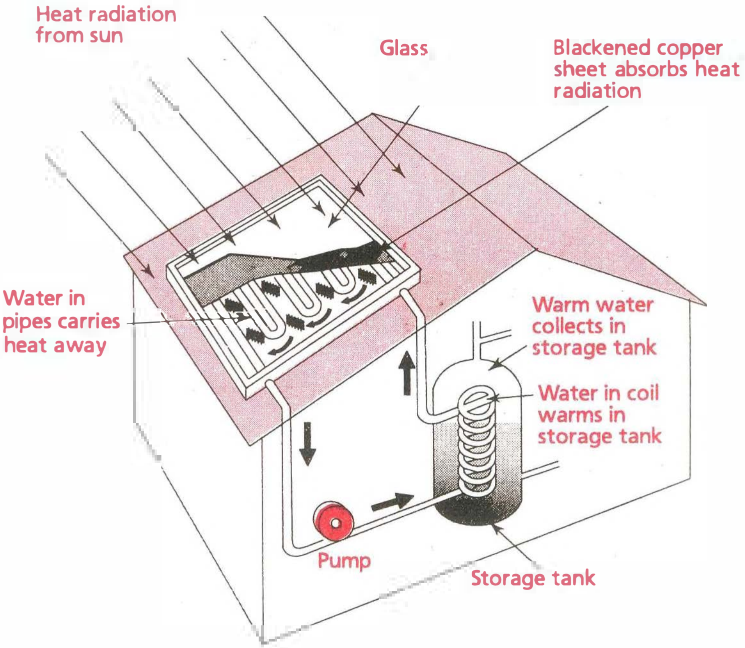

Heat radiation from sun
Glass
Blackened copper sheet absorbs heat radiation
Water in pipes carries heat away
Warm water collects in storage tank
Water in coil warms in storage tank
Pump
Storage tank

Fig. 4.1 Solar panel

Huge *solar furnaces* use the sun's heat to make electricity. A field of giant mirrors are used to collect the sun's rays and concentrate it onto a furnace where the heat boils water to make steam. The steam drives a turbine to make electricity.

Solar heaters, solar cookers, also work on solar energy.

Do you know ?

There is enough nuclear energy left in the Sun to keep it shining for another 5000 million years. On an average, each square metre of the earth's surface receives the same energy from the Sun as it would from a one-bar electrical fire.

Hydro energy

The energy in fast moving rivers or waterfalls is usually being wasted. It can be harnessed to provide a source of power, which can be used to produce electricity. The resulting *hydroelectric power* can be carried wherever it is wanted. Usually, it is necessary to build a large dam to hold back the water in a reservoir. Large pipes then carry the water from the reservoir to turbines, which drive the generators to produce electricity (Fig. 4.2).

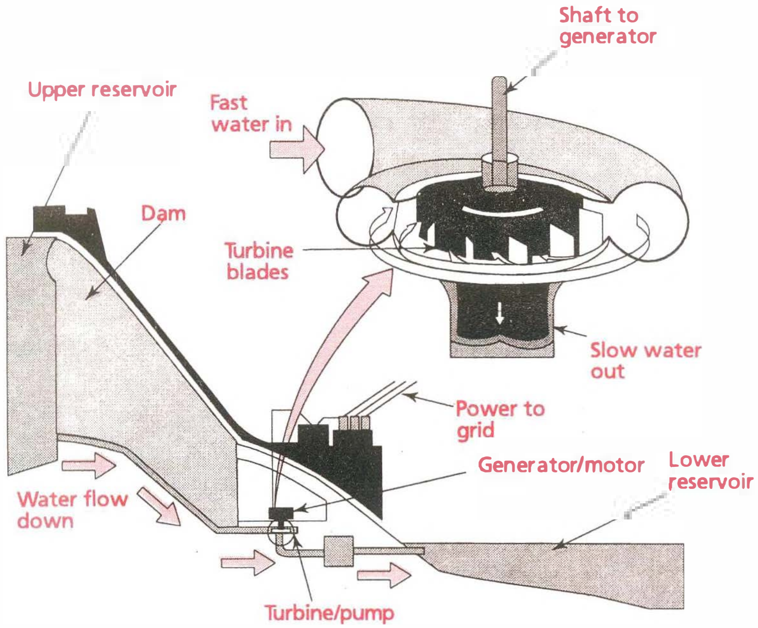


Fig. 4.2 Working of a turbine to generate electricity

In Canada, New Zealand and other countries with large rivers and waterfalls, much electricity is generated from water power. An important application of hydroelectricity is *pumped storage.* In a pumped storage system, pipes run between two reservoirs at different levels. When water is allowed to run down from the higher to the lower level, it passes through a turbine and generates electricity.

Wind energy

The energy produced from fast moving winds is called *wind energy.* It can also be harnessed to provide a source of power. Sailing ships and yachts have sailed round the world on wind power alone. Wind mills use the wind power for grinding corn and pumping water. Today, aerogenerators are using wind energy to generate electricity. Aerogenerators are placed on exposed, windy sites, often in large groups called *wind farms*. Unlike fuel-burning power stations, they do not pollute the atmosphere. But, as the force of the wind is unpredictable, only a few sites in the world are suitable.

Tidal energy

Along every coast, the tide comes in and out twice a day as the sea-level rises and falls. The movement of water can be used as a source of energy to generate electricity. This is called *tidal power*.

In a tidal power scheme, a dam is built across the mouth of a river. When the tide comes in, water rushes through tunnels in the dam and fills a basin behind. When the tides goes out again, the basin empties and water rushes back through the tunnels. In the tunnels, a flow of water in either direction turns huge turbines. These drive generators to produce electricity (Fig. 4.3).

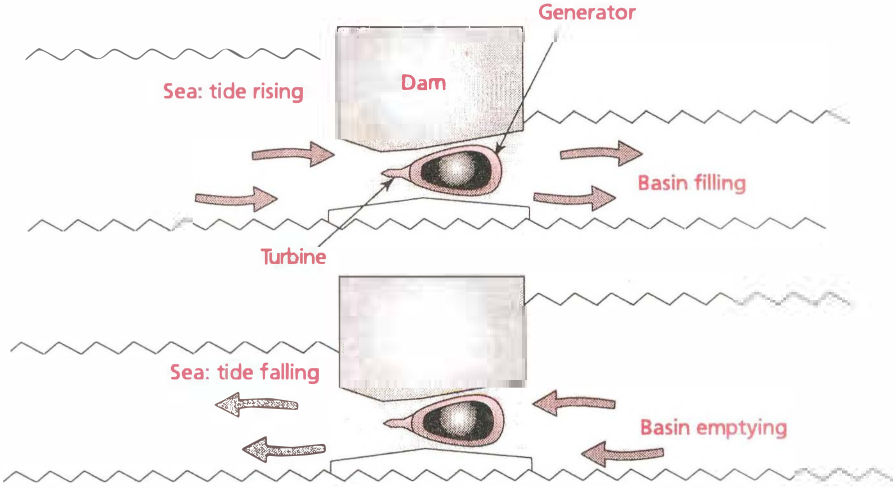


Fig. 4.3 Working of a tidal barrage

Tidal power stations burn no fuel, so there is no chance of pollution. However, they are expensive to build and can damage the local environment and the wildlife. Also, the tides occur at different times each day, so power cannot always be generated when it is most needed.

Do you know ?

The word's first major tidal power station was opened in 1966, on the Rance Estuary, Brittany, France. It has 24 generators with a total power output of 240,000 kilowatts—about a tenth of the output of a large fuel burning power station.

Geothermal Energy

Deep in the earth, the rock is hot. The deeper one goes, the hotter it gets. The heat comes from the molten rock that lies below the earth's crust. This heat energy in the interior of the earth is called *geothermal energy*. This could be used to heat water for homes and industry, and to make steam for electricity (Fig. 4.4). Geothermal energy could meet about 10 % of our present electricity needs.

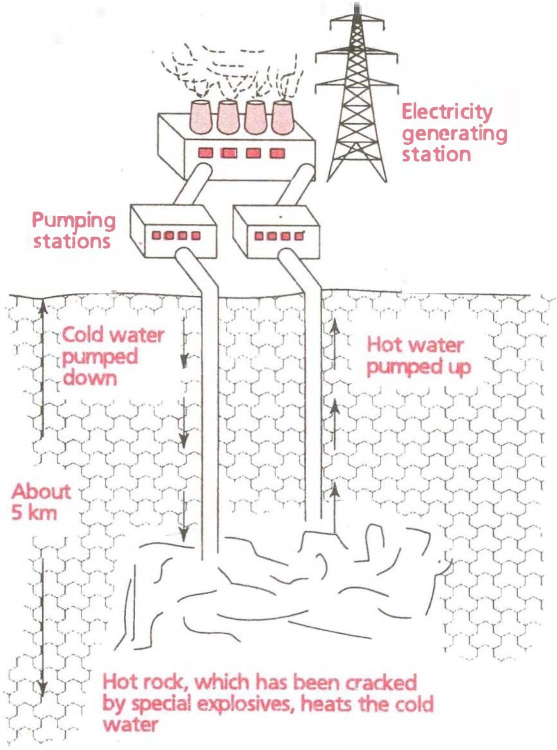

Electricity generating station
Pumping stations
Cold water pumped down
Hot water pumped up
About 5 km
Hot rock, which has been cracked by special explosives, heats the cold water

Fig. 4.4 Hot rocks used to heat water

Bio-energy

It is obtained by the decomposition of plants and animals in the absence of air. Decomposition of garbage, cowdung, sewage and other plant residues produce methane and carbon dioxide. This mixture is called *biogas*. It is used as a manure and as a fuel for cooking, running engines and street lighting.

In a sewage digester, sewage is broken down by bacteria to form methane, which is then used to fuel the sewage works. The same process can be used to make biogas from rotting vegetation.

NON-RENEWABLE ENERGY SOURCES

Sources of energy which are accumulated in nature and get exhausted with use are called non-renewable sources of energy.

They are also called *conventional sources* of energy, eg, fossil fuels, nuclear energy, etc.

Fossil fuels

Coal, petroleum and natural gas are *fossil fuels*, ie, they are the remains of plants and animals that lived millions of years ago. They produce carbon dioxide, water vapour and a plenty of energy when burnt. They are used to generate electricity, which is in turn used for heating, lighting and driving machinery (Fig. 4.5).

Nuclear energy

When radioactive elements like uranium, thorium, undergo nuclear splitting, nuclear energy is liberated. When uranium atom is bombarded with neutrons, its nucleus splits

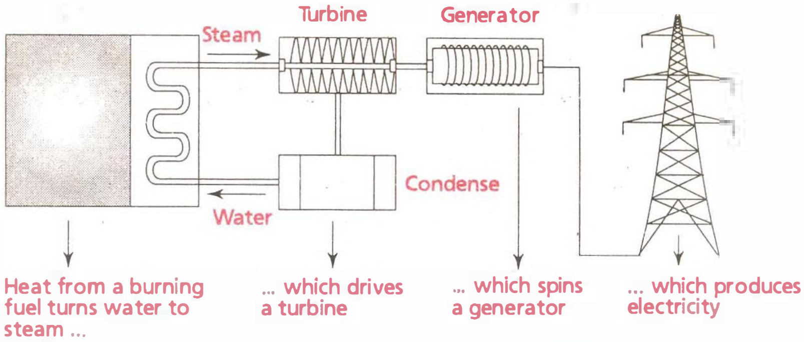

Turbine
Generator
Steam
Condense
Water
Heat from a burning fuel turns water to steam ...
... which drives a turbine
... which spins a generator
... which produces electricity

Fig. 4.5 Fuels generate electricity

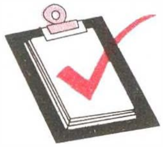
to yield atoms of smaller elements. This reaction is called *nuclear fission*. If this process is not controlled, it results in a chain reaction. However, fission reactions are controlled in special devices called *nuclear reactors*.

The heat liberated from these reactors is used to convert water to steam which is in turn, used to drive a turbine that generates electricity. Nuclear energy obtained from nuclear reactors could provide a much more efficient source of energy compared to conventional fuels. One tonne of enriched uranium nuclear fuel can produce as much energy as 25,000 tonnes of coal.

How long will these non-renewable sources of energy last?

Fossil fuels will not last forever. It is estimated that the world's supply of oil will only last for another 30 to 60 years, and coal for another 250 years. Fossil fuels will eventually run out. Even the uranium for nuclear power stations is expected to run out in another 30 years. So, there is an urgent need to develop something to replace them. The alternative sources of energy from sun, wind, tides, earth, biogas and water could prove beneficial.

Let's Answer . . .



1. Classify the following as renewable and non-renewable sources of energy:

   Sun, coal, petroleum, wind, water, uranium atom, hot rocks, biogas.
2. Why do we need to depend on renewable sources of energy?

CONSERVATION OF ENERGY

Judicious and economic use of energy resources is called conservation of energy.

Some of the ways in which energy can be conserved are:

- Reducing too much dependence on non-renewable sources of energy. More use of water and wind energy to generate electricity; use of electricity or sunlight to run motor vehicles rather than petrol or diesel, could help save our energy resources.

- Use of biogas for cooking and other domestic uses. Sugarcane can be turned into alcohol which can then be used as engine fuel.
- Use of devices based on solar energy like solar cells, cookers, etc.
- Switching of electrical instruments when not in use.
- Use of pressure cookers for cooking saves fuel.

Remember . . .

- Sun, wind, water, tides, earth, biogas, are renewable sources of energy. They are produced in nature continuously and do not get exhausted with use.
- Fossil fuels and nuclear energy are non-renewable sources of energy. They get exhausted with continuous usage.
- Judicious and economic use of energy resources is called conservation of energy.

Questions

Answer in brief

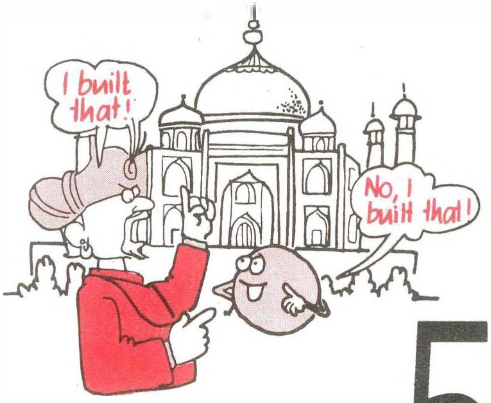
1. What are renewable sources of energy? Give examples.
2. Explain how solar panels could help save energy in the home.
3. How do solar cells turn sunlight directly into electricity?
4. How can tidal power be used to generate electricity?
5. What is geothermal energy? Explain how we can make use of it.
6. Give two examples of non-renewable energy sources and explain them briefly.
7. What are the different ways in which energy can be conserved?
8. How can animal waste be used as a source of energy?

Fill in the blanks

1. Coal, petroleum and natural gas are called _____ fuels.
2. Solar cells convert _____ into electricity.
3. Wind mills work on _____ power.
4. _____ is used as a manure and fuel for cooking and running engines.
5. _____ and _____ use the sun's heat to make electricity.

Think & Answer . . .

1. Coal is a finite energy source. What does that mean?
2. Name some more ways in which energy can be conserved in your homes and outside.


I built that!
No, I built that!

5

rocks, minerals & metals

- Rocks
- Minerals
- Metals
- Metallurgy

ROCKS

Rocks are substances that make up the Earth. We see rocks everywhere –sandstone, granite and marble, in the walls of large buildings, the cliffs along the seashore, and the broken pieces of stone in the tarmac on the roads. Rocks vary from place to place–in their colour, texture and strength.

The study of rocks present in the earth's crust is called *geology*, and the scientists who study about rocks and their behaviour are called *geologists*.

There are eight elements which make up more than 98% of the world's rocks. They are–oxygen, silicon, aluminium, iron, calcium, sodium, potassium and magnesium.

On the basis of their mode of formation, rocks are classified into three main types–*igneous rocks*, *sedimentary rocks* and *metamorphic rocks*.

Igneous rocks The oldest rocks that existed deep inside the earth's crust (extreme temperature and pressure conditions) were the molten rocks, which were called *magma*. As a result of volcanic eruptions, the magma made its way to the earth's surface, where it slowed down on cooling and ultimately solidified into hard rocks called *igneous rocks* (the lava). The kind of igneous rock formed depends on the composition of the magma and the rate at which it cools.

Examples of igneous rocks are basalt, granite, lava, gabbro and pumice. Metals like copper, iron, tin, and aluminium are also obtained from these rocks.

Sedimentary rocks Under the influence of climatic variations like wind, water, frost and snow, the rocks get eroded and are swept down the rivers and streams into the sea. Here, they settle down in layers and gradually from *sedimentary rocks*. The act of sedimentation is visible near the banks of rivers and on sea beaches as ripples, furrows and footprints.

Sedimentary rocks often contain *fossils*. Fossils are the imprints of dead plant and animal bodies which inhabited the earth several million years ago, when the rocks were formed (Fig. 5.1).

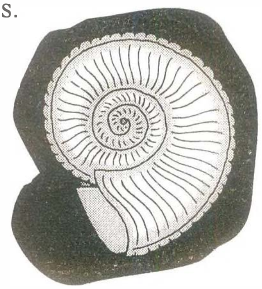


Fig. 5.1 Fossil in a piece of limestone

Examples of sedimentary rocks are *limestone* (calcium carbonate), *sandstone* (silicon oxide and calcium carbonate), *clay* (aluminium silicate), *coal*, *salt* (sodium chloride), *calcite* (calcium carbonate), coral, flint, shale and travertine.

Metamorphic rocks The word 'metamorphic' means *later, or changed form.* These are rocks which have been altered either by heat, or by heat and pressure together. Gneiss, marble, quartzite, slate and soapstone are some examples of metamorphic rocks.

MINERALS

Minerals are the chemical ingredients of rocks. They include everyday substances such as rocksalt, asbestos, graphite used in pencil lead, talc used in talcum powders, china clay used to make crockery, etc. They are mostly solids and have a definite chemical composition.

More than 2500 minerals are known to us. They can be identified by studying the distinctive properties of their individual components, ie, colour, lustre, hardness, specific gravity, etc.

Minerals may be *metallic* or *non-metallic.* Examples of metallic minerals are iron, copper, gold and silver. Non-metallic minerals include quartz, mica, rock salt, etc. Minerals like petroleum, coal, which originate from the remains of dead plants and animals are called *organic minerals.*

Many minerals were formed from *magma* which is the molten material of the earth's interior. When magma hardens underground, it cools very slowly. The chemicals in it crystallize to form minerals such as mica, quartz and feldspar.

Do you know ?

In the Moh's scale, minerals are arranged in increasing order of their hardness. This scale was devised by the German scientist, Friedrich Mohs in the year 1773-1839. Any mineral will scratch one that is softer than itself but not one that is harder.

1. Talc *Softest*
2. Gypsum
3. Calcite
4. Fluorite
5. Apatite
6. Feldspar
7. Quartz
8. Topaz
9. Corundum
10. Diamond *Hardest*

Table 5.1 Some minerals and their uses

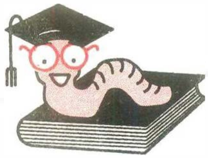
Minerals	Uses
1. Magnetite (Fe_3O_4)	Source of iron.
2. Limestone ($CaCO_3$)	As a building material and as a raw material for cement, soda glass, iron and steel. Used as a source of cheap alkali, eg, lime is used to neutralize acidity in soil and in refining sugar.

3. Quartz (SiO_2)	Glass, sand paper, telephones, radio, watches.
4. Silver	Jewellery, coins, etc.
5. Mica	As an insulator in electrical appliances.
6. Pitch blende	As a source of uranium and radium, which are nuclear fuels.
7. Gypsum ($CaSO_4\ 2H_2O$)	Raw material for cement and plaster of paris.
8. Pyrites	Source of sulphur, iron and copper.
9. Tinstone (SnO_2)	Source of tin metal.
10. Gold and diamond	Jewellery.
11. Petroleum and coal	Fuels.

Table 5.2 Mineral wealth of India

Mineral	Location	Uses
1. Iron as haematite (Fe_2O_3)	Bihar, West Bengal, Orissa, Madhya Pradesh, Tamil Nadu and Karnataka	Used in steel manufacture.
2. Copper as copper pyrites ($CuFeS_2$)	Bihar, Rajasthan, Madhya Pradesh, Andhra Pradesh	Used in coins, utensils, statue making, electrical wires, etc.
3. Gold	Karnataka (Kolar fields), Andhra Pradesh and Tamil Nadu	Jewellery.
4. Aluminium as bauxite ($Al_2O_3\ 2H_2O$)	Jammu and Kashmir Gujarat, Bihar, Orissa, Tamil Nadu, Madhya Pradesh	Used in making utensils, packing foils, parts of aeroplanes, ships, cars, trains, etc, and overhead electrical wires.
5. Limestone ($CaCO_3$)	Rajasthan and Madhya Pradesh	As a building material.
6. Chromium as chromite ($FeCr_2O_4$)	Bihar, Orissa, Maharashtra and Karnataka	Production of special steels and for plating of metals.
7. Titanium as ilmenite ($FeTiO_3$)	Kerala and Tamil Nadu	Used in fabrication of jet engines, marine equipments, supersonic aircrafts, hardening of steel, etc.
8. Manganese as pyrolusite (MnO_2)	Madhya Pradesh, Orissa, Bihar, Gujarat, Rajasthan, Karnataka, and Andhra Pradesh	Production of manganese steel. Its alloys are used in making parts of ships, etc.
9. Mica	Bihar, Orissa, Tamil Nadu and Rajasthan. India is the largest producer of Mica.	Used as an electrical insulator.
10. Thorium as monazite.	Kerala	As a future nuclear fuel. Its oxide compound is used to prepare incandescent gas mantles.
11. Coal	Bihar, West Bengal, Tamil Nadu	Used as a fuel.
12. Petroleum	Mumbai, Gujarat, Assam, and recently in offshore deltas of Cauvery, Godavari and Krishna.	Used as a fuel.

Remember . . .



- Rocks are classified into three types—igneous, metamorphic and sedimentary rocks, based on their mode of formation.
- Minerals are the chemical ingredients of rocks. They are naturally occurring elements or compounds.

METALS

The earth's crust contains about 90 different elements. More than 60 of these elements are metals such as gold, silver, copper, tin and iron. Metals occur in nature both in the free and combined states. The naturally occurring compounds of metals are the minerals. Most minerals have some impurities like clay, sand, etc. The impurities present in minerals are called the *gangue* (matrix).

Metals cannot be extracted from all minerals. The minerals from which metals are extracted are called the *ores*. Ores are oxides, sulphides, sulphates, carbonates and other compounds of metals. All ores are minerals but all minerals are not ores.

Table 5.3 Relative abundance of metals in the earth's crust

Metal	Percentage abundance	Metal	Percentage abundance
Potassium	1.5 %	Calcium	5 %
Sodium	2.5 %	Magnesium	3 %
Aluminium	8 %	Other metals	2 %
Iron	6 %		

Aluminium is the most abundant of all metals in the earth's crust.

Table 5.4 Common ores and metals extracted from them

Ore	Chemical Formula	Metals Extracted
Bauxite	$Al_2O_3.2H_2O$	Al
Corundum	Al_2O_3	Al
Haematite	Fe_2O_3	Fe
Magnetite	Fe_3O_4	Fe
Zincite	ZnO	Zn
Tin stone	SnO_2	Sn
Cuprite	Cu_2O	Cu
Pyrolusite	MnO_2	Mn
Monazite	ThO_2	Th
Galena	PbS	Pb
Copper pyrites	$CuFeS_2$	Cu
Copper glance	CuS	Cu
Iron pyrites	FeS	Fe

(Table 5.4 *Contd*)

(**Table 5.4** *Contd*)

Zinc blende	ZnS	Zn
Cinnabar	HgS	Hg
Dolomite	$CaCO_3.MgCO_3$	Mg
Magnesite	$MgCO_3$	Mg
Malachite	$Cu(OH)_2\ CuCO_3$	Cu
Siderite	$FeCO_3$	Fe
Calamine	$ZnCO_3$	Zn
Cerussite	$PbCO_3$	Pb
Limestone, chalk, marble	$CaCO_3$	Ca
Rocksalt	$NaCl$	Na
Horn silver	$AgCl$	Ag
Carnallite	$KCl.MgCl_2\ 6H_2O$	Mg
Fluorspar	CaF_2	Ca
Cryolite	$AlF_3.3NaF$	Al

METALLURGY

The extraction of a metal from the ore depends on the type of ore being used, nature of impurities present and reactivity of the metal to be extracted.

The various steps involved in the extraction of metals from their ores, and their refining is known as metallurgy.

The various steps in metallurgy are:

(i) Concentration of the ore.

(ii) Conversion of the concentrated ore into metal oxide.

(iii) Reduction of metal oxide to metal.

(iv) Refining of impure metal.

Concentration of the ore

Ore is an impure source of the metal. By removing the gangue or matrix, a concentrated ore is obtained with higher percentage of the metal. Concentration begins with *crushing* the ore in big jaw crushers and then *grinding* it in a ball mill to get powdered ore. Concentration is also known as *dressing* of the ore. Some common methods of concentration are:

(i) Hydraulic washing

(ii) Froth floatation

(iii) Magnetic separation.

Hydraulic washing This method of concentration is based on the difference in the densities of the ore and gangue. A stream of water is made to pass through the powdered ore kept on a long wooden table with obstacles called cleats. The heavier ore particles are obstructed by the cleats while the lighter impurities pass over along with the stream of water. Ores of *tin* and *lead* are concentrated by this method (Fig. 5.2).

Froth floatation This method is used to concentrate sulphide ores. It is based on the principle of preferential wetting of the ore and gangue by certain liquids. The ore is wetted by certain oils like pine oil and the gangue is wetted by water.

In this process, the sulphide ore is taken in a tank full of water to which small quantities of oil are added (Fig. 5.3). The water is agitated violently by blowing in compressed air to from *froth* at the water interface. The ore particles are left behind in the water. The froth separates out and is collected in the other tank. On standing, the froth settles down and concentrated ore is obtained. This process is used to concentrate sulphide ores of *copper*, *lead* and *zinc*.

Magnetic separation This method is based on the principle of difference in the magnetic properties of the ore and the gangue. A magnetic separator consists of a brass or leather belt moving on two rollers, of which one is magnetic (Fig. 5.4).

The powdered ore is dropped over the belt from one end. The magnetic contents of the ore are attracted towards the magnetic roller at the other end. The non-magnetic impurities fall away from the magnetic roller. Ores of *iron*, *nickel* and *manganese* are concentrated by this method.

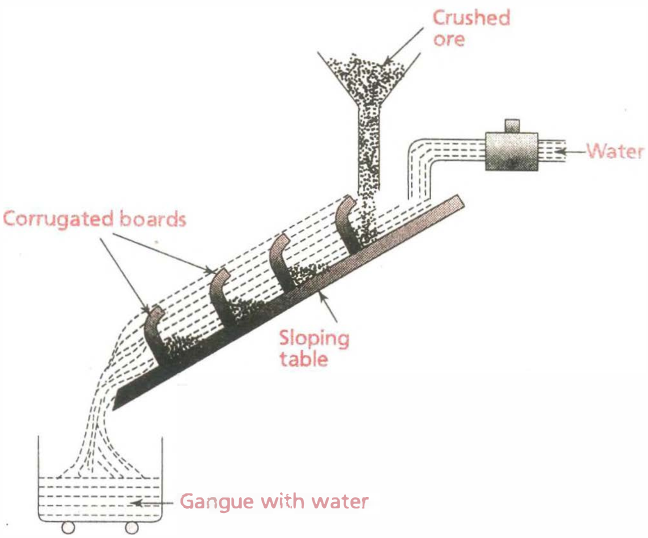


Fig. 5.2 Hydraulic washing of the ore

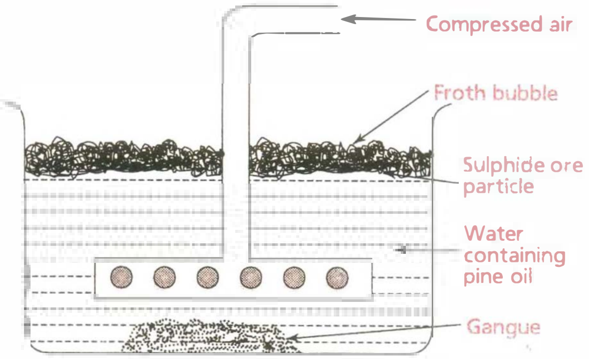


Fig. 5.3 Concentration of the ore by froth floatation

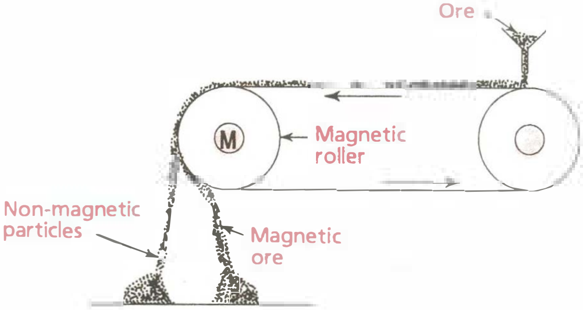


Fig. 5.4 Magnetic separation of the ore

Conversion of the concentrated ore into metal oxide

The concentrated ore is converted into metal oxide by *calcination* or *roasting*. This method depends on the nature of the ore used. A *carbonate* ore is subjected to calcination and a *sulphide* ore is roasted.

Calcination In this process, the carbonate ore is heated strongly in the absence of air. It gets converted into the metal oxide. Water and other volatile impurities present in the ore are removed.

For example:

$$\underset{\text{(calamine)}}{ZnCO_3} \xrightarrow{\Delta} ZnO + CO_2\uparrow$$

$$\underset{\text{(limestone)}}{CaCO_3} \xrightarrow{\Delta} CaO + CO_2\uparrow$$

Roasting In this process, the sulphide ore is heated strongly in the presence of air. It is oxidized to the metal oxide. Moisture and other volatile impurities are removed.

For example:

$$\underset{\text{(zinc blende)}}{2ZnS} + 3O_2 \xrightarrow{\Delta} 2ZnO + 2SO_2$$

Reduction of metal oxide to metal

The process of converting a metal oxide into a metal is known as *reduction*. When the process is carried out using carbon and aluminium, it is known as *chemical reduction*.

For example:

$$ZnO + C \longrightarrow Zn + CO$$

$$3MnO_2 + 4Al \longrightarrow 3Mn + 2Al_2O_3$$

Refining of the impure metal

Refining of the metal is done to remove the physical and chemical impurities present in it. It depends on the nature of metal extracted and the nature of impurities present.

Refining by electrolysis is called *electrolytic refining*. This method is used for refining metals like *copper*, *silver*, *tin*, *zinc*, *nickel*, *chromium* and *gold*. The apparatus used for electrolytic refining is shown in figure 5.5. A solution of the soluble salt of a metal known as the *electrolyte* is taken in a tank. A thick block of impure metal is made the *anode* (positive electrode) and a thin sheet of pure metal is made the *cathode* (negative electrode). When electric current is passed through the electrolyte, the pure metal gets deposited at the cathode. All the impurities settle down below the anode as *anode mud* in the solution.

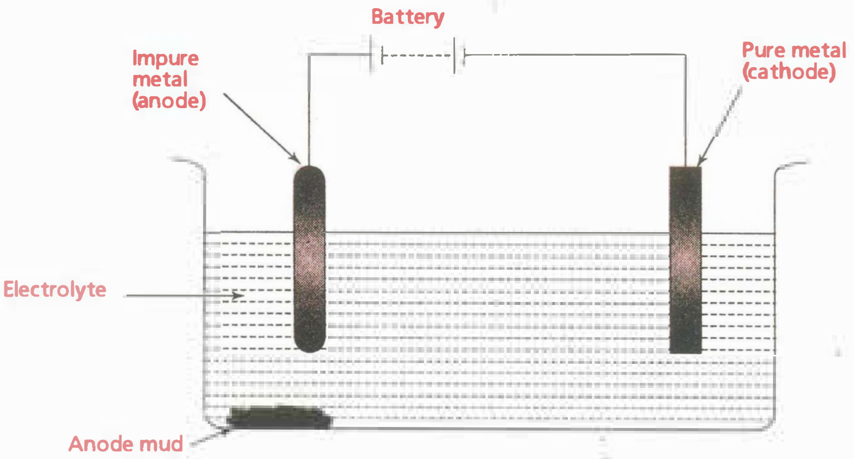


Fig. 5.5 Electrolytic refining

Other methods of refining include *liquation*, *distillation* and *oxidation*.

Let's Answer . . .

1. Which process is used to concentrate
   (a) sulphide ore of lead?
   (b) oxide ore of tin?
   (c) oxide ore of iron?
2. Which method is used to convert
   (a) sulphide ore of zinc into its oxide, in the presence of air?
   (b) carbonate ore of zinc into its oxide, in the absence of air?
3. Which property of iron is used to separate it from other ore impurities?
4. Name the rock from which metals are extracted.

Remember . . .

- All ores are minerals but all minerals are not ores.
- The various steps involved in the extraction of metals from their ores and its refining is called metallurgy.
- Concentration of the ore is done by (i) hydraulic washing (ii) froth floatation and (iii) magnetic separation.
- Calcination or roasting is done to convert a concentrated ore into its metal oxide.
- Reduction of metal oxide to the metal is achieved by reducing agents like C and Al.
- Electrolytic refining is widely used for the purification of metals.

Questions

Answer in brief

1. How are rocks classified? Give examples.

2. Define a mineral and mention any three uses of it.
3. What are ores? List the various steps involved in the extraction of a metal from its ore?
4. Write a note on (a) froth floatation (b) hydraulic washing (c) magnetic separation (d) electrolytic refining.
5. Differentiate (a) ore and mineral (b) calcination and roasting.

Fill in the blanks

1. Study of rocks is called _____.
2. The impurities associated with a mineral are called _____.
3. Granite is a _____rock.
4. Sulphide ores are concentrated by _____.
5. Hydraulic washing is done to concentrate ores of _____ and _____.

Choose the correct alternative

1. Limestone and sandstone are
   (a) igneous rocks (b) sedimentary rocks (c) metamorphic rocks
2. Gneiss is an example of
   (a) ore (b) mineral (c) metamorphic rock
3. Bauxite is the ore of
   (a) sodium (b) aluminium (c) tin
4. Haematite is a
   (a) sulphide ore (b) carbonate ore (c) oxide ore
5. Galena is
   (a) ZnS (b) PbS (c) CuS

Think & Answer . . .

1. Take a sulphide ore of zinc as sample. Starting from concentration of the ore to extraction of metal in the pure form, write the various steps involved. Also write equations where necessary.
2. An oxide ore of iron contains sand, clay and copper as impurities. How would you concentrate this ore?
3. A metal X is extracted from its sulphide ore. Which would be the most suitable method for concentrating this ore?

extraction of metals 6

- Extraction
- Iron • Copper
- Aluminium • Alloys

In the previous chapter, we have seen that metals are obtained from ores. An ore is usually a compound of the metal, mixed with impurities. When the ore has been dug up, it must be decomposed in some way, to get the metal. This is called *extracting* the metal. The method of extraction depends on how reactive the metal is. The more reactive the metal, the more difficult it is to decompose its compounds.

Table 6.1 shows the common methods of extracting metals from their ores.

Table 6.1 Common methods of extracting metals

Metal	Method of extraction from ore	
Potassium	Electrolysis of the hydroxide	Most reactive
Sodium	Electrolysis of the chloride	
Calcium	Electrolysis of the chloride	
Magnesium	Electrolysis of the chloride	
Aluminium	Electrolysis of the oxide	
Copper	Roasting in air	
Silver, gold	Occur naturally as elements	Least reactive

Electrolysis is the most powerful way to decompose a metal ore. But it needs a lot of electricity, which makes it expensive.

Let us study the extraction of a few metals in detail.

IRON

Iron is present to the extent of 5 % in the earth's crust. It occurs in the native state in minor quantities. Unreacted iron is also present in small quantities in meteorites. The common ores of iron are–

(i) Haematite – Fe_2O_3
(ii) Magnetite – Fe_3O_4
(iii) Limonite – $Fe_2O_3\,3H_2O$
(iv) Siderite – $FeCO_3$
(v) Iron pyrites – FeS.

Iron is chiefly extracted from haematite ore by smelting it in a blast furnace (Fig. 6.1).

- The blast furnace is fed with a *furnace charge* which consists of iron ore, limestone (calcium carbonate), coke (made from coal) and hot air.
- Mixture of iron ore, limestone and coke are crushed and fed into the top of the blast furnace.
- Hot air is blown at the bottom of the furnace at 800 °C through tubes called *tuyers.* The coke is oxidized to carbon dioxide.

$$\underset{\text{(coke)}}{C} + O_2 \longrightarrow \underset{\text{(carbon dioxide)}}{CO_2} + \text{Heat}$$

- The carbon dioxide formed is further reduced by coke to form carbon monoxide.

$$CO_2 + C \longrightarrow \underset{\text{(carbon monoxide)}}{2CO}$$

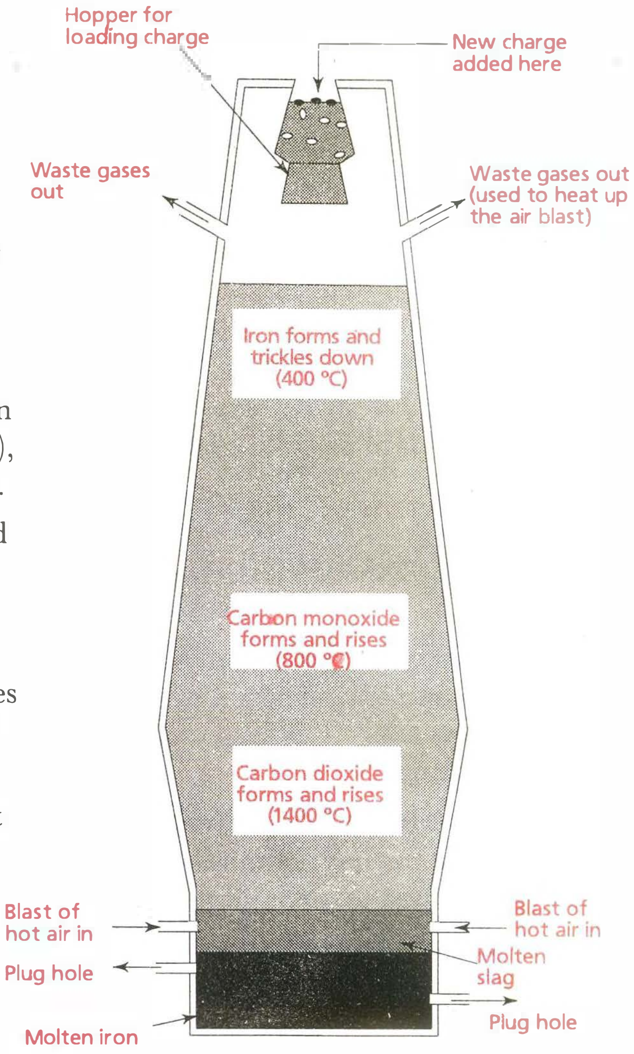

Hopper for loading charge
New charge added here
Waste gases out
Waste gases out (used to heat up the air blast)
Iron forms and trickles down (400 °C)
Carbon monoxide forms and rises (800 °C)
Carbon dioxide forms and rises (1400 °C)
Blast of hot air in
Blast of hot air in
Plug hole
Molten slag
Molten iron
Plug hole

Fig. 6.1 Extraction of iron by blast furnace process

- Carbon monoxide is a *reducing agent* and it reduces the iron (III) oxide to molten iron. The iron so produced is in the liquid form which trickles down to the bottom of the furnace. The molten iron is run off from the bottom into moulds where it solidifies on cooling.

$$3CO + Fe_2O_3 \longrightarrow 2Fe + 3CO_2$$

Function of limestone ($CaCO_3$)

Due to the action of heat, limestone decomposes to calcium oxide and carbon dioxide.

$$\underset{\text{(limestone)}}{CaCO_3} \xrightarrow{\Delta} \underset{\text{(calcium oxide)}}{CaO} + CO_2$$

The silica impurity of the iron ore reacts with the calcium oxide obtained from limestone to form molten calcium silicate, referred to as *slag*.

$$CaO + \underset{\text{(flux)}}{SiO_2} \longrightarrow \underset{\text{(slag)}}{CaSiO_3}$$

Molten slag floats over molten iron at the bottom of the furnace.

As extraction proceeds, more and more limestone, ore, and coke are fed into the top of the furnace and the molten slag and iron are tapped off at the bottom.

The iron obtained from the blast furnace is called *cast iron or pig iron.* It is brittle due to the presence of large amounts of carbon and other dissolved impurities.

Commercial forms of iron

There are three different commercial forms of iron. They are *cast iron*, *wrought iron* and *steel.*

Cast iron Cast iron is the cheapest form of iron. It contains 2.5–5 % carbon and other impurities like sulphur, silicon, phosphorus, etc. It is brittle and has a low tensile strength. It rusts easily. It is used to make articles like manhole lids, radiators, railings, etc, and in making steel.

Wrought iron It is the purest form of iron containing about 0.2 % carbon. It is malleable and ductile but has low tensile strength than steel. It can be welded and does not rust easily. It is used to make cores of electromagnets, iron nails, pipes, bolts, gates, etc.

Steel Steel is an alloy of iron with 0.1–1.5 % of carbon. It is prepared by mixing molten iron with scrap iron and limestone. The charge is heated in an oxygen converter and oxygen is blown into the molten mixture. Limestone converts the impurities into slag and most of the carbon escapes as carbon dioxide.

Table 6.2 Different types of iron—A Comparison

Nature/ Properties/ Uses	Cast iron	Wrought iron	Steel
1. Purity	Impure	Almost pure	Not pure but commercially important.
2. Impurities	Silicon, phosphorus, sulphur, manganese.	Traces–less than 0.5 %	Carbon and other added metals.
3. Carbon content	2.5–5 %	0.1–0.25 %	0.1–1.5 %

(Table 6.2 *Contd*)

(**Table 6.2** *Contd*)

4.	Hardness	Hard	Soft	Hardness depends on % carbon.
5.	Brittleness	Brittle	Not brittle	Depends on carbon content.
6.	Malleability	Non-malleable	Malleable	Malleable
7.	Ductility	Non-Ductile	Ductile	Ductile
8.	Corrosion	Resists corrosion	Resists corrosion	Ordinary steel corrodes.
9.	Effect on cooling	Expands	Does not expand	Does not expand
10.	Melting point	1200 °C	1500 °C	1200 °C–1500 °C
11.	Uses	Weight, drains, pipes, railings, radiators.	Chains, nails, anchors, electro-magnets.	Automobile parts, blades, ships, train parts, rails, buildings.

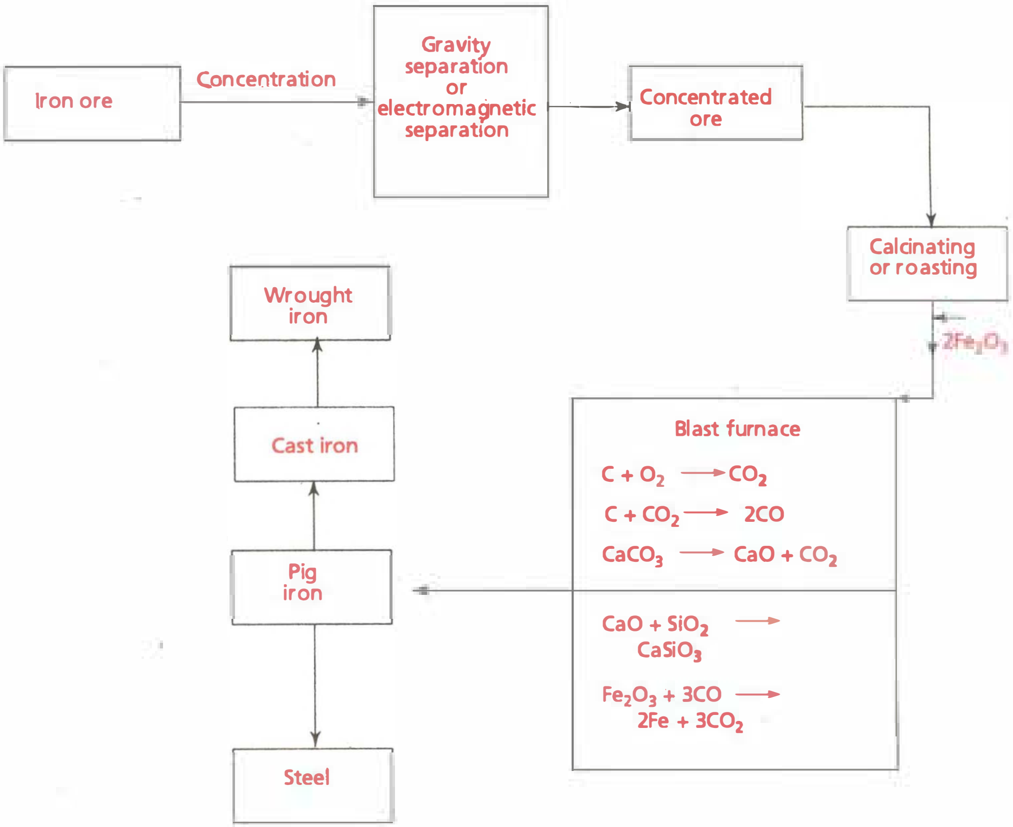

Iron ore
Concentration
Gravity separation or electromagnetic separation
Concentrated ore
Calcinating or roasting
$2Fe_2O_3$
Blast furnace
$C + O_2 \longrightarrow CO_2$
$C + CO_2 \longrightarrow 2CO$
$CaCO_3 \longrightarrow CaO + CO_2$
$CaO + SiO_2 \longrightarrow CaSiO_3$
$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$
Pig iron
Cast iron
Wrought iron
Steel

Fig. 6.2 Flow chart for the extraction of iron

COPPER

The important ores of copper are–

- (i) Copper pyrites – $CuFeS_2$
- (ii) Cuprite – Cu_2O
- (iii) Malachite – $CuCO_3\,Cu(OH)_2$
- (iv) Copper glance – Cu_2S

Copper is chiefly extracted from copper pyrites which also contains small quantities of iron. The various steps involved in the extraction of copper metal from copper pyrites are:

- The ore is crushed and ground to a fine powder. It is then concentrated by froth floatation process to remove the gangue particles.
- The concentrated pyrite ore is roasted in excess of air. A mixture of copper sulphide and iron sulphide is obtained. The impurities present in the ore get oxidized.

$$\underset{\text{(copper pyrite)}}{2CuFeS_2} + O_2 \xrightarrow{\text{roasting}} Cu_2S + 2FeS + SO_2$$

- The roasted ore ($Cu_2S + FeS$) is mixed with some coke and sand (silica). The mixture is smelted in a furnace.
- Silica acts as a flux and reacts with iron oxide to form ferrous silicate (slag).

$$FeO + \underset{\text{(flux)}}{SiO_2} \longrightarrow \underset{\text{(slag)}}{FeSiO_3}$$

- The slag is poured off leaving behind a molten mixture of Cu_2S and a little FeS, which is called *matte.*
- The matte is transferred into another furnace called *bessemer convertor* and a blast of hot air mixed with sand is blown in. Some of the copper sulphide reacts with oxygen to form copper oxide.

$$2Cu_2S + 3O_2 \longrightarrow 2Cu_2O + 2SO_2$$

- The copper oxide reacts with the remaining copper sulphide to form copper metal.

$$2Cu_2O + Cu_2S \longrightarrow 6Cu + SO_2.$$

- The copper obtained is called *blister copper.* It contains 95 % copper and 5 % impurities.
- Blister copper is purified by electrolytic refining. In this process, impure copper is dipped into a solution of copper sulphate (electrolyte) and made to form the anode. A strip of pure copper is also dipped in the electrolyte and acts as the cathode. Electric current is passed through the solution. Impure copper dissolves at the anode and

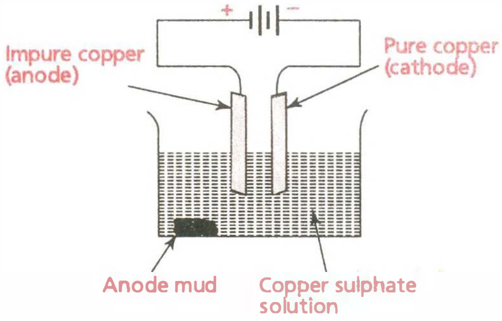


Fig. 6.3 Electrolytic refining of copper

pure copper gets deposited at the cathode. The impurities collect below the anode as anode mud (Fig. 6.3).

Uses of copper

1. Copper is used for making electric wires, as it is a good conductor of electricity.
2. It is used in making utensils, coins, containers, etc.
3. It is used in making alloys such as brass, bronze and German silver.
4. It is used in electroplating.
5. It is used in the manufacture of calorimeters, vacuum pumps and printed electronic circuits.

ALUMINIUM

Aluminium is the most abundant element in the earth's crust. The important ores of aluminium are –

(i) Bauxite – $Al_2O_3\ 2H_2O$
(ii) Cryolite – Na_3AlF_6
(iii) Corundum – Al_2O_3
(iv) Kaolin – $Al_2Si_2O_7\ 2H_2O$

Aluminium is chiefly extracted from bauxite ore.

- Crude bauxite is crushed, ground, dried and heated with concentrated caustic soda solution (NaOH). Bauxite dissolves in NaOH forming sodium aluminate solution. The undissolved impurities are filtered off.

$$\underset{\text{(bauxite)}}{Al_2O_3} + 2NaOH \longrightarrow \underset{\text{(sodium aluminate)}}{2NaAlO_2} + H_2O$$

- Sodium aluminate solution is diluted with water and treated with a little aluminium hydroxide to cause precipitate formation. Hydrolysis takes place and a white precipitate of $Al(OH)_3$ is formed.

$$NaAlO_2 + 2H_2O \xrightarrow{\text{hydrolysis}} \underset{\text{(white ppt.)}}{Al(OH)_3} + NaOH$$

- The aluminium hydroxide precipitate is filtered, washed, then heated to get pure aluminium oxide (alumina).

$$2Al(OH)_3 \xrightarrow{\Delta} \underset{\text{(alumina)}}{Al_2O_3} + 3H_2O$$

- *Electrolytic reduction of alumina*

Aluminium is extracted from alumina by electrolysis. Alumina (Al_2O_3) is mixed with a calculated quantity of cryolite (Na_3AlF_6) and melted in an iron cell. Pure alumina melts at 2045°C. So, cryolite is added to lower its melting point.

The iron cell is lined with carbon, which acts as the cathode. Graphite rods act as the anode. Electric current is passed through the circuit (Fig. 6.4). Molten aluminium is liberated at the cathode and sinks to the bottom of the cell from where it is removed. Oxygen gas is liberated at the anode.

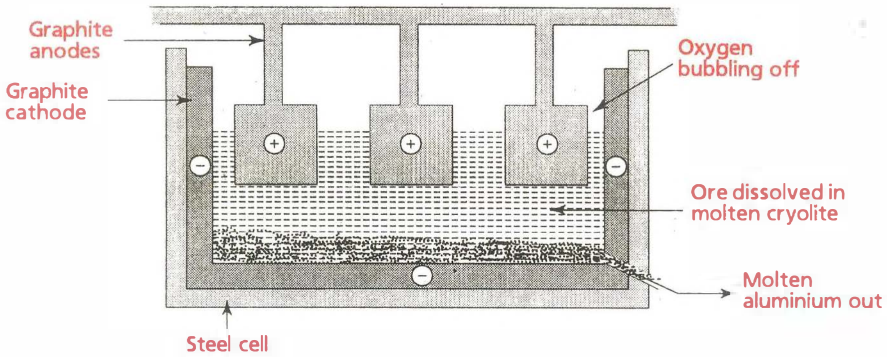


Fig. 6.4 Electrolytic reduction of alumina

$$2Al_2O_3 \xrightarrow{\text{electrolysis}} 4Al + 3O_2$$

As the metal is removed, fresh amounts of pure alumina are added. The anode has to be replaced from time to time, because oxygen gas attacks the graphite anodes and eats them away.

Uses of aluminium

1. It is used for making electrical wires and cables.
2. It is used as a reducing agent in metallurgy.
3. It is used for making cooking utensils, foils, saucepans, TV aerials, ships, aeroplanes and even space rockets.
4. It is used for making silver paints and mirrors.
5. It is used in making alloys, such as duralumin and magnalium, which are used for making aircraft bodies.

Do you know

Aluminium and iron are involved in a competition reaction to repair rails. Aluminium and iron (III) oxide are being heated together, to give molten iron. This is run into gaps between railway lines. The process is called *Thermit process*.

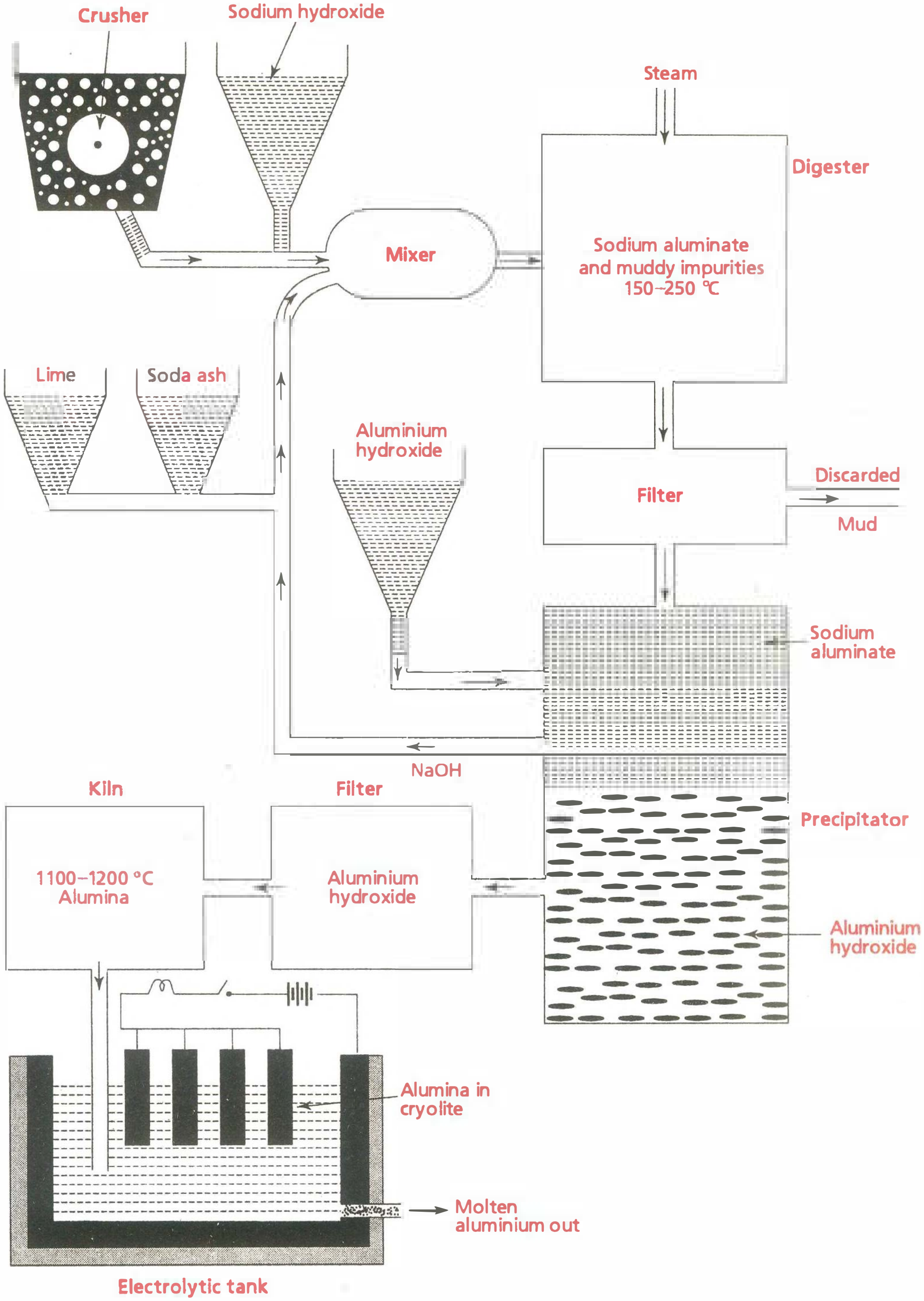


Fig. 6.5 Flow chart for the extraction of aluminium

Let's Answer . . .

1. Electrolysis is needed for extracting calcium. Why is it not needed for extracting copper?
2. In the electrolysis of alumina, why must the anodes be replaced from time to time?
3. Complete:

   The chief ore of aluminium is _____. It is first purified to give _____ or _____, whose formula is _____. This is _____ to give pure aluminium.

Remember . . .

- Most metals are extracted from their ores.
- The more reactive the metal, the more powerful and expensive the extraction.
- Iron is mainly extracted from haematite ore by smelting it in a blast furnace
- Copper is chiefly extracted from copper pyrites ore.
- Aluminium is extracted from bauxite ore by electrolysis.

ALLOYS

Some metals are useful when they are pure. For example, pure copper is easily drawn into wires and is used in electrical circuits, since it is a good conductor of electricity. Pure aluminium can be rolled into very thin sheets, which are quite strong but can be cut easily. It can be used for milk tops and cooking foils.

But many metals are more useful when they are not pure. One such example is iron. Pure iron is not good for building purposes, since it is too soft and stretches easily. Besides, it also rusts. But when a little carbon (0.5%) is mixed with it, it becomes harder and stronger. This mixture is called *mild steel.* It is used for buildings, bridges, ships and car bodies. Similarly, when nickel and chromium are mixed with iron, it becomes harder and no longer rusts. It is called *stainless steel,* and is used for car parts, utensils, etc. Thus, we see that the properties of iron have been changed, by mixing other substances with it. Such mixtures are called *alloys.*

Alloy is a homogeneous mixture of a metal with other metals or non-metals.

Alloys are usually made by melting the main metal and then dissolving other substances in it. Some common alloys, their composition and uses are summarized in table 6.3.

Table 6.3 Alloys—composition and uses

S. No.	Alloys	Contents	Uses
1.	Brass	Copper and zinc	For making utensils, flower vases, fancy electrical fittings, screws, bolts, pipes, scientific instruments and jewellery.
2.	Bronze	Copper and tin	For making coins, statues, medals, cooking utensils, ship propellers, etc.
3.	German silver	Copper, zinc and nickel	For making utensils, silver wares and in electroplating.
4.	Solder	Lead and tin	For soldering purposes in electronic industry.
5.	Steel	Iron and carbon	For making ships, vehicles, screws, nuts, bridges, etc.
6.	Stainless steel	Iron, chromium and nickel	Cycles and automobile parts, tools, cooking utensils, in dairy industry.
7.	Invar	Iron and nickel	For making clock pendulums and measuring instruments.
8.	Alnico	Aluminium, nickel, cobalt.	For making magnets.
9.	Duralumin	Aluminium, copper, magnesium, manganese	For making pressure cooker, aircrafts and body parts of satellites.
10.	Magnalium	Aluminium and magnesium	Making light instruments.

Remember . . .

- Alloy is a homogeneous mixture of a metal with other metals or non-metals.
- Alloys are usually made by melting the main metal and then dissolving other substances in it.

Questions

Answer in brief

1. Name any three ores of copper, iron and aluminium.
2. How is iron extracted from haematite?
3. Write the various steps involved in the extraction of copper from copper pyrites.
4. How is aluminium obtained from alumina?
5. What are alloys? Name any three alloys of copper and give their uses.

6. What are the different forms of iron? Explain.
7. How is iron more useful when it is mixed with other metals?
8. Write equations for the following reactions:
   (a) Carbon monoxide reduces ferric oxide to molten iron.
   (b) Copper pyrite ore is roasted in excess of air.
   (c) Bauxite is treated with concentrated caustic soda solution.

Fill in the blanks

1. German silver is an alloy of _____, _____ and _____.
2. _____ is added to lower the melting point of alumina.
3. _____ is the purest form of iron.
4. Matte is a mixture of _____ and _____.
5. Bronze is an alloy of _____ and _____.
6. The formula of corundum is _____.
7. Limonite is an ore of _____.
8. More reactive metals are extracted by _____.

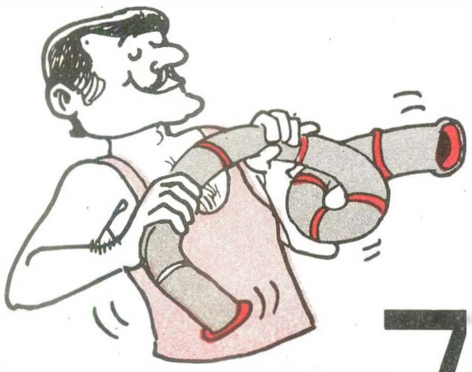
Choose the correct alternative

1. In the extraction of iron, silica is used as
   (a) flux (b) ore (c) slag (d) reducing agent
2. Siderite is an ore of
   (a) iron (b) copper (c) tin (d) aluminium
3. Brass is an alloy of
   (a) copper and zinc (b) copper and tin
   (c) copper and lead (d) none of these
4. Copper pyrites is concentrated by
   (a) roasting (b) froth floatation (c) calcination (d) electrolysis
5. During electrolysis, impure metal
   (a) acts as cathode (b) acts as anode (c) acts as electrolyte
   (d) none of these.

Think & Answer . . .

1. Explain why
   (a) aluminium is used for window frames.
   (b) iron is used in bridges.
   (c) copper is used for electrical wiring.
2. Iron and aluminium ores are relatively cheap, but aluminium metal is a lot more expensive than iron metal. Why?

3. A metal X is extracted from its ore Y. The properties of the metal are given below.
   (a) The metal is bluish-white and silvery.
   (b) It is the most abundant element in the earth's crust.
   (c) It is obtained by electrolysis.
   Identify 'X'. Also give the possible formula for Y.



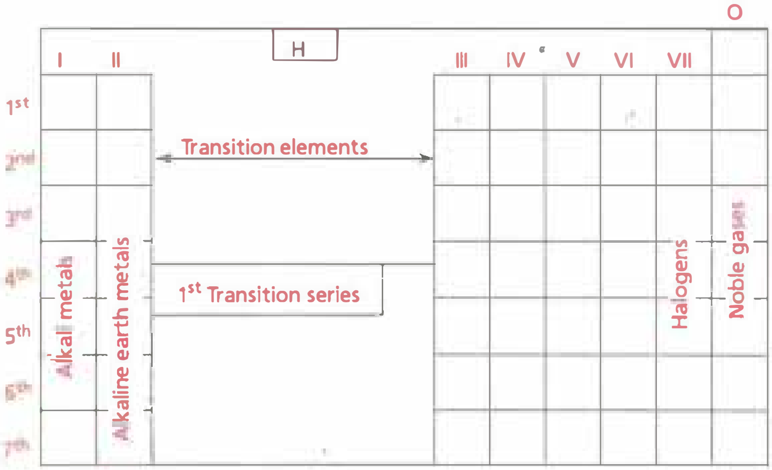
7

properties of metals

- Physical Properties of Metals
- Chemical Properties of Metals
- Comparison of Metals and Non-metals
- Corrosion

You have already learnt in lower classes that elements are pure substances comprising one type of atoms only. More than 100 elements are known to us, out of which 92 are naturally occurring and the rest are produced in laboratories. Since the early nineteenth century, many attempts have been made to find some pattern among the chemical elements. The most successful attempt was that by the Russian chemist, Mendeleev, who arranged the elements in a periodic table (Table 7.1).

Table 7.1 Periodic Table



In this table, Mendeleev

(a) arranged elements in order of increasing atomic mass.

(b) grouped elements with similar chemical properties in columns.

(c) left gaps in the table for missing elements. He predicted the existence and properties of these elements. Later, the elements *germanium* and *gallium* were discovered and Mendeleev's predictions were shown to be correct.

In the modern periodic table, Mendeleev arranged chemical elements in order of increasing atomic numbers. The horizontal rows of elements are called *periods* and the vertical columns of elements are called *groups*. The modern Periodic Table is shown in table 7.2. In this table, elements lying to the left of the zig zag line (below it) are *metals* and to the right (above it) are *non-metals*. You will study more about Periodic Table in higher classes.

PHYSICAL PROPERTIES OF METALS

Metals differ widely in their properties. But all of them have certain common properties.

Metals are solids All metals are solids, except mercury which exists in liquid form.

Metals possess metallic lustre Metals in pure state shine. This property is called *metallic lustre*. The shiny surface on metals makes them attractive for decorative purposes and making jewellery, eg, gold and silver are used in jewellery and copper is used for decorative purposes. A thin coating of silver is used at the back of mirrors since it is a good reflector of light.

Metals are generally hard Metals like iron, copper, gold, silver and aluminium are hard and strong. They cannot be cut by a knife. But there are some metals like sodium, potassium and magnesium that are soft and can be cut by a knife.

Metals are malleable and ductile Metals are *malleable*, ie, they can be hammered into sheets without breaking, eg, gold and silver. Copper, aluminium, tin, iron can be beaten to very thin sheets called foils. Metals are *ductile*, ie, they can be drawn into thin wires, eg, gold, silver, copper and aluminium.

Metals are good conductors of heat and electricity Metals like copper, aluminium, are good conductors of heat and electricity. *Silver* is the best conductor of heat and electricity. The poorest conductor of heat is *lead* and the poorest conductor of electricity are *iron* and *mercury*.

The electrical conductivity of metals increases in the order

$$\mathrm{Pb} < \mathrm{Fe} < \mathrm{Zn} < \mathrm{Al} < \mathrm{Au} < \mathrm{Cu} < \mathrm{Ag}$$

Metals are sonorous They produce a metallic sound when struck.

Metals have high melting and boiling points Iron melts at 1535 °C. The melting points of copper, aluminium, zinc and lead are 1083 °C, 660 °C, 420 °C and 327 °C, respectively. Mercury is a liquid metal, which freezes at −39 °C.

Metals have high densities Density varies from metal to metal. Metals like gold, mercury, silver, iron, have high densities. Some metals like sodium, potassium, magnesium have low densities. Lithium is the lightest metal (density 0.5 g/cm^3).

Table 7.2 The Modern Periodic Table

Non-metals

Metals

Transition elements

Group I	II											Group III	IV	V	VI	VII	0
1 1 **H** Hydrogen																	4 2 **He** Helium
7 3 **Li** Lithium	9 4 **Be** Beryllium											11 5 **B** Boron	12 6 **C** Carbon	14 7 **N** Nitrogen	16 8 **O** Oxygen	19 9 **F** Fluorine	20 10 **Ne** Neon
23 11 **Na** Sodium	24 12 **Mg** Magnesium											27 13 **Al** Aluminium	28 14 **Si** Silicon	31 15 **P** Phosphorus	32 16 **S** Sulphur	35.5 17 **Cl** Chlorine	40 18 **Ar** Argon
39 19 **K** Potassium	40 20 **Ca** Calcium	45 21 **Sc** Scandium	48 22 **Ti** Titanium	51 23 **V** Vanadium	52 24 **Cr** Chromium	56 25 **Mn** Manganese	57 26 **Fe** Iron	58 27 **Co** Cobalt	59 28 **Ni** Nickel	64 29 **Cu** Copper	65 30 **Zn** Zinc	70 31 **Ga** Gallium	73 32 **Ge** Germanium	75 33 **As** Arsenic	79 34 **Se** Selenium	80 35 **Br** Bromine	84 36 **Kr** Krypton
85 37 **Rb** Rubidium	88 38 **Sr** Strontium	89 39 **Y** Yttrium	91 40 **Zr** Zirconium	93 41 **Nb** Niobium	96 42 **Mo** Molybdenum	98 43 **Tc** Technetium	100 44 **Ru** Ruthenium	102 45 **Rh** Rhodium	106 46 **Pd** Palladium	108 47 **Ag** Silver	112 48 **Cd** Cadmium	115 49 **In** Indium	119 50 **Sn** Tin	122 51 **Sb** Antimony	128 52 **Te** Tellurium	127 53 **I** Iodine	131 54 **Xe** Xenon
133 55 **Cs** Caesium	137 56 **Ba** Barium	139 57 ***La** Lanthanum	178.5 72 **Hf** Hafnium	181 73 **Ta** Tantalum	184 74 **W** Tungsten	186 75 **Re** Rhenium	190 76 **Os** Osmium	192 77 **Ir** Iridium	195 78 **Pt** Platinum	197 79 **Au** Gold	201 80 **Hg** Mercury	204 81 **Tl** Thallium	207 82 **Pb** Lead	209 83 **Bi** Bismuth	210 84 **Po** Polonium	210 85 **At** Astatine	222 86 **Rn** Radon
223 87 **Fr** Francium	226 88 **Ra** Radium	227 89 **⊕Ac** Actinium															

	140 58 **Ce** Cerium	141 59 **Pr** Praseodymium	144 60 **Nd** Neodymium	147 61 **Pm** Promethium	150 62 **Sm** Samarium	152 63 **Eu** Europium	157 64 **Gd** Gadolinium	159 65 **Tb** Terbium	162 66 **Dy** Dysprosium	165 67 **Ho** Holmium	167 68 **Er** Erbium	169 69 **Tm** Thulium	173 70 **Yb** Ytterbium	175 71 **Lu** Lutetium
⊕	232 90 **Th** Thorium	231 91 **Pa** Protactinium	238 92 **U** Uranium	237 93 **Np** Neptunium	242 94 **Pu** Plutonium	243 95 **Am** Americium	247 96 **Cm** Curium	247 97 **Bk** Berkelium	251 98 **Cf** Californium	254 99 **Es** Einsteinium	253 100 **Fm** Fermium	256 101 **Md** Mendelevium	254 102 **No** Nobelium	257 103 **Lr** Lawrencium

Let's Answer . . .

1. Give two reasons why metals are used in drums and in saucepans?
2. Why is silver used in jewellery and in mirrors?
3. Why is aluminium used in cooking foils?

Remember . . .

- In the modern periodic table, elements (metals and non-metals) are arranged in increasing order of atomic numbers.
- Metals are hard, lustrous, malleable, ductile and good conductors of heat and electricity.
- Metals have high densities and melting/boiling points.

CHEMICAL PROPERTIES OF METALS

Reaction with oxygen

Metals react with oxygen to form *metal oxides.* These oxides are *basic* in nature. They turn red litmus solution blue.

$$\text{Metal} + \text{Oxygen} \longrightarrow \text{Oxide}$$

Metals like *sodium* and *potassium* react vigorously with oxygen at room temperature, to form sodium oxide and potassium oxide.

$$4Na + O_2 \longrightarrow \underset{\text{(sodium oxide)}}{2Na_2O}$$

$$4K + O_2 \longrightarrow \underset{\text{(potassium oxide)}}{2K_2O}$$

Magnesium burns brightly in air to form magnesium oxide.

$$2Mg + O_2 \longrightarrow \underset{\text{(magnesium oxide)}}{2MgO}$$

Zinc and *lead* react with oxygen to form a protective coating of an oxide layer.

$$2Zn + O_2 \longrightarrow \underset{\text{(zinc oxide)}}{2ZnO}$$

$$2Pb + O_2 \longrightarrow \underset{\text{(lead oxide)}}{2PbO}$$

Copper reacts with oxygen to form copper oxide which gets deposited on the surface of the metal.

$$2Cu + O_2 \longrightarrow \underset{\text{(copper oxide)}}{2CuO}$$

Gold, platinum and *silver* do not react with oxygen. They are called *noble metals.*

iron reacts with oxygen to form a brown layer of iron oxide (rust) on its surface. This reaction is known as *rusting of iron.*

Table 7.3 Reaction of some metals with oxygen

Metal	Behaviour	Order of reactivity	Product
Sodium	Catches fire with only a little heating. Burns fiercely with a bright yellow flame.	Most reactive	Sodium oxide (Na_2O). Also forms sodium peroxide (Na_2O_2), a pale yellow powder.
Magnesium	Catches fire easily. Burns with a white flame.		Magnesium oxide (MgO), a white powder.
Iron	Does not burn, but the hot metal glows brightly in oxygen, and gives off yellow sparks.		Iron oxide (Fe_3O_4), a brownish-black powder.
Copper	Does not burn, but the hot metal gets coated with a black substance.		Copper oxide (CuO), a black powder
Gold	No reaction	Least reactive	

Reaction with water

Metals react with cold water or steam depending on their position in the reactivity series (Fig. 7.1).

Sodium, *potassium* and *calcium* react vigorously with water at room temperature to form an alkaline solution of a hydroxide with the liberation of hydrogen gas.

$$2Na + 2H_2O \longrightarrow 2NaOH + H_2\uparrow$$

$$2K + 2H_2O \longrightarrow 2KOH + H_2\uparrow$$

$$Ca + 2H_2O \longrightarrow Ca(OH)_2 + H_2\uparrow$$

Magnesium reacts slowly with cold water but violently with steam to form magnesium oxide and hydrogen.

$$Mg + H_2O \longrightarrow MgO + H_2\uparrow$$

Aluminium, *zinc* and *iron* react very slowly with steam. *Gold*, *copper* and *silver* do not react with cold water or steam.

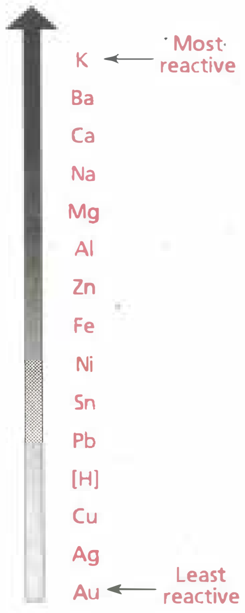


Fig. 7.1 Reactivity series of metals

Table 7.4 Reaction of metals with water

Metal	Behaviour	Order of reactivity	Product
Potassium	Violent with cold water. Catches fire.	Most reactive	KOH solution and hydrogen gas.
Sodium	Violent with cold water.		NaOH solution and hydrogen gas.
Calcium	Less violent with cold water.		$Ca(OH)_2$ solution (slightly soluble) and hydrogen.
Magnesium	Very slow with cold water, but vigorous with steam		Solid MgO and hydrogen.
Zinc	Quite slow with steam.		Solid ZnO and hydrogen.
Iron	Slow with steam		Solid Fe_3O_4 and hydrogen.
Gold, copper, silver	No reaction	Least reactive	—

Activity 1

Aim To study the reaction of magnesium, with steam.

Materials required A boiling tube, bunsen burner, wool soaked in water, a coil of magnesium ribbon, tongs.

Method Place the wet wool at the bottom of a boiling tube. Insert a magnesium coil into the tube and heat it over a burner. Magnesium burns brightly as it reacts with steam. A white powder of magnesium oxide is left behind. A colourless gas is given out that burns with a pop sound. it is hydrogen (Fig. 7.2).

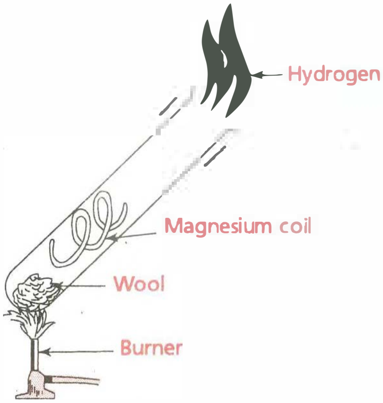


Fig. 7.2 Action of steam on magnesium

$$Mg + \underset{\text{(steam)}}{H_2O} \longrightarrow MgO + H_2\uparrow$$

Reaction with acids

Most metals react with dilute acids to produce salt and hydrogen.

$$\text{Metal} + \text{Acid(dil)} \longrightarrow \text{Salt} + \text{Hydrogen}$$

Sodium and *potassium* react vigorously with dilute hydrochloric acid, with the liberation of tremendous amount of heat.

$$2Na + 2HCl \longrightarrow 2NaCl + H_2\uparrow$$
$$2K + 2HCl \longrightarrow 2KCl + H_2\uparrow$$

Magnesium, calcium, zinc and *aluminium* also react with dilute hydrochloric acid to liberate hydrogen.

$$Mg + 2HCl \longrightarrow MgCl_2 + H_2\uparrow$$
$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2\uparrow$$
$$2Al + 6HCl \longrightarrow 2AlCl_3 + 3H_2\uparrow$$

Iron reacts very slowly with dilute hydrochloric acid.

$$Fe + 2HCl \longrightarrow FeCl_2 + H_2\uparrow$$

Lead does not react with dilute acids. It reacts very slowly with concentrated hydrochloric acid to form a solution of lead chloride and hydrogen.

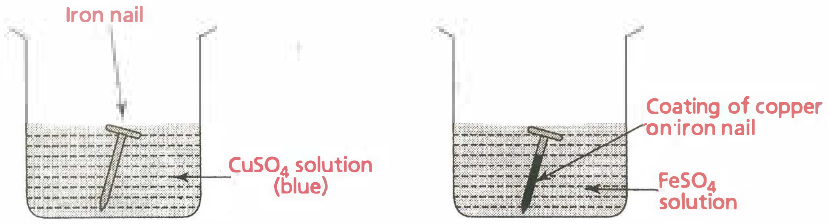
$$Pb + 2HCl \longrightarrow PbCl_2 + H_2\uparrow$$

Copper, silver and *gold* do not react with dilute or concentrated acids.

All metals above hydrogen in the reactivity series displace hydrogen from dilute acids.

Displacement reaction

Consider a reaction involving iron and copper (Fig. 7.3).



(a) Iron nail dipped in $CuSO_4$ solution

(b) The solution turns pale green due to the formation of $FeSO_4$

Fig. 7.3

$$\underset{\text{(iron)}}{Fe} + \underset{\text{(copper sulphate)}}{CuSO_4} \longrightarrow \underset{\text{(iron sulphate)}}{FeSO_4} + \underset{\text{(copper)}}{Cu}$$

Iron is more reactive than copper and displaces copper from its salt solution.

A more reactive metal will always displace a less reactive metal from its salt solution.

A few more displacement reactions are given below:

$$Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu \text{ [Zn is more reactive than Cu]}$$
$$Cu + ZnSO_4 \longrightarrow \text{No reaction [Cu is less reactive than Zn]}$$
$$Mg + CuSO_4 \longrightarrow MgSO_4 + Cu \text{ [Mg is more reactive than Cu]}$$

$$Cu + MgSO_4 \longrightarrow \text{No reaction [Cu is less reactive than Mg]}$$

$$Fe + CuSO_4 \longrightarrow FeSO_4 + Cu \text{ [Iron is more reactive than Cu]}$$

$$Cu + FeSO_4 \longrightarrow \text{No reaction [Cu is less reactive than Fe]}$$

Activity 2

Aim To show that copper is more reactive than silver.

Materials required Copper coil, silver nitrate solution, test tube.

Method Half fill a test tube with silver nitrate solution ($AgNO_3$). Dip a coil of copper into the solution. Observe for sometime. The copper gradually dissolves and the solution turns blue. The copper coil gets coated with silver (Fig. 7.4). What does this show?

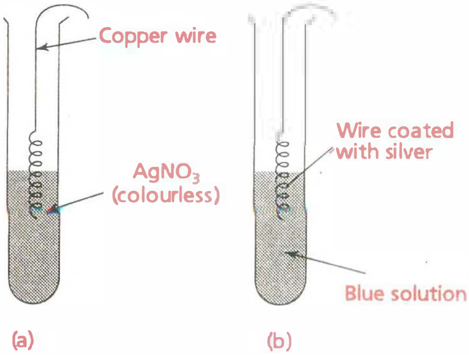


Fig. 7.4

Copper is more reactive than silver. Hence it displaces silver from its salt solution

$$\underset{\text{(copper)}}{Cu} + \underset{\text{(silver nitrate)}}{2AgNO_3} \longrightarrow \underset{\text{(copper nitrate)}}{Cu(NO_3)_2} + \underset{\text{(silver)}}{2Ag}$$

Let's Answer . . .

1. What happens when zinc is added to copper sulphate solution? Write the equation.
2. Tin does not react with iron (III) oxide. But it reduces lead (II) oxide to lead. Arrange tin, iron and lead in order of decreasing reactivity.
3. Will copper react with lead (II) oxide? Explain.

Remember . . .

- Metals react with oxygen to form metal oxide.
- Metals react with cold water or steam to form hydroxides, or oxides with the liberation of hydrogen.
- All metals above hydrogen in the reactivity series displace hydrogen from dilute acids.
- A more reactive metal can displace a less reactive metal from its salt solution.

COMPARISON OF METALS AND NON-METALS

Metals and non-metals differ in many aspects. Table 7.4 shows the differences between metals and non-metals.

Table 7.4 Metals and non-metals—A comparison

	Metals	Non-metals
1.	They are lustrous and can be polished.	They are dull and cannot be polished (exceptions–iodine and graphite).
2.	They are malleable and ductile.	They are not malleable or ductile.
3.	They are good conductors of heat and electricity.	They are poor conductors of heat and electricity (except graphite).
4.	They are mostly solids (except–mercury).	They are mostly solids or gases (except bromine).
5.	They can be easily cut or bent	They are brittle.
6.	They have high melting and boiling points.	They have low melting and boiling points.
7.	They are sonorous.	They are not sonorous.
8.	They have high densities.	They have low densities.
9.	They react with oxygen to form basic oxides.	They react with oxygen to form acidic oxides.
10.	Examples–gold, silver copper, lead, sodium, etc.	Examples–sulphur, carbon, chlorine, bromine, iodine, etc.

CORROSION

Some metals *corrode* when exposed to air. For example, iron articles develop a brown coating called rust on their surface when exposed to moist air. Aluminium vessels lose their shine and copper articles turn green when exposed to damp air.

The process of oxidation of metals under the influence of atmospheric conditions is known as corrosion.

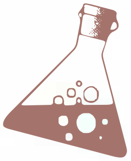
Any chemical or electrochemical reaction on the surface of a metal is also known as corrosion.

Some metals like gold and platinum do not corrode in air. They are used in making coins and jewellery.

Rusting Corrosion of iron and steel is called *rusting*. Rust is hydrated Iron (III) oxide, ie, $Fe_2O_3 \cdot 3H_2O$. Rusting takes place in the presence of air (oxygen) and water.

Rusting is an electrochemical process in which different parts of the iron surface act as electrodes in a cell reaction. It is accelerated by the presence of impurities in the iron and acids or other electrolytes in water.

Activity 3



Aim To show that both air and water are necessary for rusting.

Materials required Three test tubes, stop cork, three nails, wax, tap water, boiled water, calcium chloride powder.

Method Take three test tubes A, B and C. Half fill test tube A with tap water, B with dry calcium chloride and C with boiled water (boiled to drive away all dissolved air). Place an iron nail in each of the three test tubes. Close the mouth of test tube B with stop cork and seal the mouth of test tube C with wax (Fig. 7.5). Leave it for a day or two. Do you observe any change? Explain what happens.

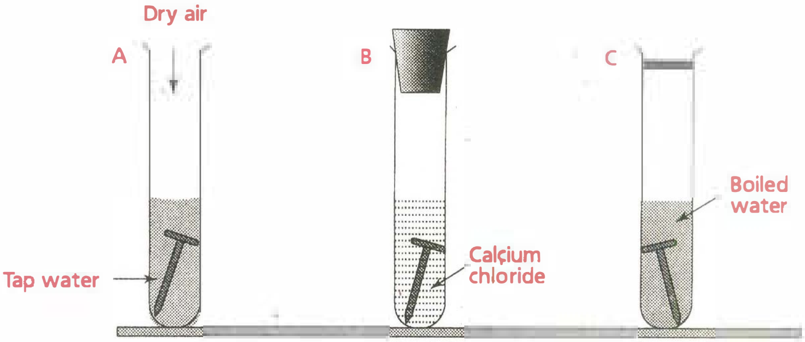


Fig. 7.5 Moisture is necessary for rusting

The iron nail in test tube A rusts while those in B and C do not. This is because the iron nail in B is not in contact with air or water and the one in C is in contact with water but not exposed to air. This shows that both air and water are necessary for rusting.

Prevention of corrosion

Corrosion can be prevented by the following methods:

Paint If the surface of the metal is covered with a coating of paint, corrosion can be prevented. For example, railway coaches, automobiles, grills, pipes, steel furniture, etc, are painted to prevent corrosion. The paints are made from zinc or any other metal which is more reactive than iron and can prevent it from reacting with the moisture in air.

Oiling and greasing Iron and steel machinery parts, tools, etc, are smeared with oil or greases, when not in use, to prevent corrosion.

Galvanization The process of coating iron articles with a thin layer of zinc is known as *galvanization*. When exposed to air, zinc forms a protective layer of zinc oxide which prevents the contact of iron with air and hence corrosion.

Tinning It is a process of giving a coating of tin (molten tin). Steel cans are coated with a thin layer of tin to protect them from corrosion. The tin layer is, in turn, plated with a coating of lacquer or plastic. If these two layers get distorted, then the iron would be exposed to air, and tin, being less reactive than iron cannot protect it from corrosion. Zinc cannot be used in tin cans carrying preserved food because it reacts with the food contents, especially if they are acidic, like fruits.

Electroplating In this process, the metal surface is coated with another metal by electrolytic means. Gold and silver are electroplated over various articles.

Anodizing In this method, metals like aluminium, copper, etc, are coated electrically with a thin film of their oxides to protect them from corrosion. Electrolysis is carried out using dilute sulphuric acid as the electrolyte. The metal to be coated is taken as the anode.

Copper turns green when exposed to air due to the formation of a protective layer of copper carbonate (verdigris). Copper is coated electrically with an oxide layer to prevent corrosion.

Alloying A metal is mixed with other corrosion resistant metal to form an alloy, eg, stainless steel is an alloy of iron which does not rust easily.

Sacrificial protection Magnesium is more reactive than iron. When a bar of magnesium is attached to the side of a steel ship, it corrodes instead of the steel. When it is completely corroded, it can be replaced by a fresh bar. This is called *sacrificial protection*, because magnesium is sacrificed to protect the steel. Zinc can also be used in the same way.

Do you know

Nitinol or 'magic metal' is an amazing material that remembers its shape. It is an alloy of nickel and titanium. Magic metal can be given a particular shape and then twisted out of shape. On heating, it snaps back into its first shape. The metal resists corrosion, and can be used for surgical implants in medical patients.

Remember . . .

- The process of oxidation of metals under the influence of atmospheric conditions is called corrosion.
- The corrosion of iron and steel is called rusting.
- Both air and water are necessary for rusting.
- Corrosion can be prevented by galvanization, tinning, electroplating, alloying, oiling, greasing, application of paints and by sacrificial protection.

Questions

Answer in brief

1. List the various physical properties of metals.
2. Tabulate the differences between metals and non-metals.
3. Explain with examples the action of metals with oxygen.
4. How do metals react with acids? Explain.
5. What is corrosion and rusting?
6. Iron that is tin-plated does not rust. Why?
7. Explain why magnesium can prevent the rusting of iron.

8. What happens when
   (a) steam reacts with magnesium?
   (b) iron reacts with dilute hydrochloric acid?
   (c) a strip of copper is dipped in silver nitrate solution?
   (d) iron is exposed to air?
   (e) sodium reacts with oxygen?
   (f) zinc dust is added to lead nitrate solution?
   (g) copper chips are added to mercuric nitrate solution?
9. From the reactivity series, select the metal that
   (a) reacts with cold water vigorously.
   (b) does not react with dilute acids.
   (c) does not displace copper from its salt solution.
   (d) does not react with dilute or concentrated acids.
   (e) rusts in the presence of air and water.
10. What are displacement reactions? Give two examples.

Fill in the blanks

1. Gold, silver and copper are _____ reactive than sodium.
2. _____ and _____ are necessary for rusting.
3. _____ is the only liquid metal known.
4. _____ is the only non-metal which conducts electricity.
5. All metals have _____ density and are _____ conductors of _____ and _____.
6. Iron is _____ reactive than copper.
7. _____ metal burns in oxygen with a bright yellow flame.
8. _____ metal reacts less violently with cold water.

Choose the correct alternative

1. Rust is
   (a) FeO (b) Fe_2O_3 (c) $Fe_2O_3\ 3H_2O$ (d) Fe_3O_4
2. Metals react with acids to liberate
   (a) oxygen (b) carbon dioxide (c) hydrogen (d) nitrogen
3. Metals react with oxygen and form
   (a) acidic oxide (b) basic oxide (c) neutral oxide (d) none of these
4. Sodium metal is
   (a) malleable (b) ductile (c) lustrous (d) all of these.
5. The element used in electrical wires is
   (a) copper (b) sodium (c) tin (d) mercury.

Think & Answer . . .

1. Why is sodium metal always stored under kerosene?
2. Give two reasons why
   (a) mercury is used in thermometers.
   (b) aluminium is used in beer cans.
3. Which metal is not malleable at room temperature?
4. Four test tubes A, B, C and D are provided to you. All the test tubes are filled with tap water. A painted nail is placed in test tube A, nail coated with vaseline is placed in test tube B, magnesium and copper wires are wound around the iron nails in test tubes C and D, respectively (Fig. 7.6).

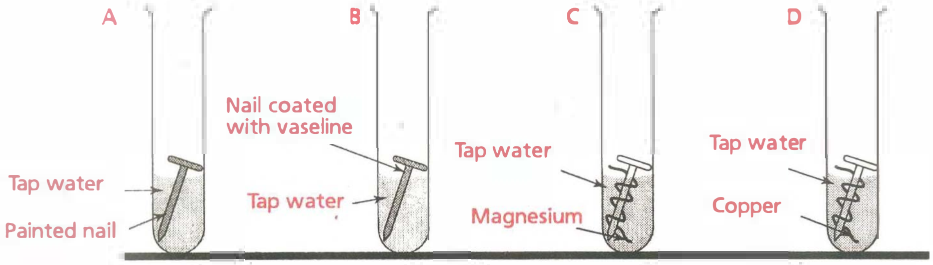


Fig. 7.6

   (a) Which of the four iron nails would get rusted?
   (b) Which of the four iron nails would remain free of rust?
   Explain your answers.

man-made materials

8

- Fibres
- Nylon
- Rayon
- Polyester
- Plastics
- Ceramics
- Soaps
- Detergents
- Fertilizers and manure
- Pesticides, fungicides and herbicides

We come across different kinds of materials, some of which are natural and some man-made. The natural materials come from plants or animals or are found in the ground, eg, wood and cork are obtained from trees; cotton from cotton plants; leather from cattle and wool from sheep; stone, slate, and gold from the ground. People could also make paper from wood, extract iron from iron ore dug out of the ground and make bricks, tiles and pottery by heating clay.

We still use most of these natural materials, either because they work well or because they look good or feel good. But we can also make use of a whole range of new materials, which are made by treating natural materials, but are better than them for many purposes. Such materials are called *man-made* or *synthetic materials.*

Substances obtained from natural materials through different chemical processes are known as man-made materials.

Concrete for building is made by mixing stone, sand, water and cement, which is itself made by heating clay and the mineral limestone. Glass for windows is made by heating sand, limestone and soda, a mineral. Steel for cars and buildings is manufactured by heating iron ore with coke and limestone. All these are examples of man-made materials.

In this chapter, we shall study in detail about a few man-made materials.

FIBRES

A *fibre* is a thread-like material that is strong and flexible enough to be converted into clothes, ropes, nets, etc. They are of two kinds–*natural* and *synthetic* fibres.

Natural fibres are long, thin threads which form parts of animals or plants. They are important to us because they are used to make fabrics.

Cotton, wool, silk, linen and *jute* are some important natural fibres. Cotton and jute are made of cellulose, which is the major constituent of the cell walls of plant cells. Wool is made from the fleece of sheeps but some of the best wools come from goats, too. Wool has the major disadvantage that it can be eaten by pests such as moths. Linen is made from the stalks of a plant called *flax*. Silk is obtained from cocoons of *silkworm*. These fibres lack resistance towards insects, moths and moulds (fungus). Cotton and linen are made up of carbon, hydrogen and oxygen only, but wool and silk contain nitrogen also.

All these materials can be spun, woven and dyed.

Synthetic fibres include *nylon*, *rayon*, *terylene*, *acrylics*, etc. Let us study about the properties and uses of a few synthetic fibres.

Do you know

The first synthetic fibre was a cellulose fibre made by the French chemist Hilaire Chardonnet in 1884. He used the fibre to produce artificial silk. Nylon was the first plastic fibre, invented by the American chemist, Wallace Carothers in 1935. Polyester and acrylics soon followed, being first developed in 1940s.

NYLON

Nylon was developed in New York and London, hence the name *nylon*. It is a polymer fibre made up of polyamide chains.

Properties

- It is very strong and elastic.
- It absorbs very little water and is thus known to possess drip-dry property.
- It is wrinkle resistant and keeps permanent creases.
- It is moth and mould resistant.
- It is light, fine and durable.

Uses

- It is used to make fishing nets, climbing ropes, tyre cords, strings for sports goods and musical instruments, bristles for tooth brushes and paint brushes, tyres.
- It is used in the production of textiles like sarees, shirts, socks, and other garments.
- It is widely used as a plastic for making machine parts.

RAYON

It is prepared from cellulose. It resembles silk in appearance and hence is also called *artificial silk*. There are different varieties of rayon, but the one prepared by *viscose process* is the most important.

Preparation of viscose rayon The raw materials used in the preparation of rayon are: cellulose $(C_6H_{10}O_5)_n$, sodium hydroxide, carbon disulphide and dilute sulphuric acid. Cellulose in the form of wood pulp is treated with NaOH and CS_2 solution to form a pale yellow syrupy liquid called *viscose*.

$$\begin{tikzcd} \text{Cellulose} \arrow[r, "1.\ \text{NaOH}", "2.\ CS_2"'] & \text{Viscose} \\ \text{wood pulp} \arrow[u] & \end{tikzcd}$$

The viscose is passed through the fine holes of a spinneret into a solution of dilute sulphuric acid, when a thread-like fibre is formed (Fig. 8.1). This product is called *viscose rayon*.

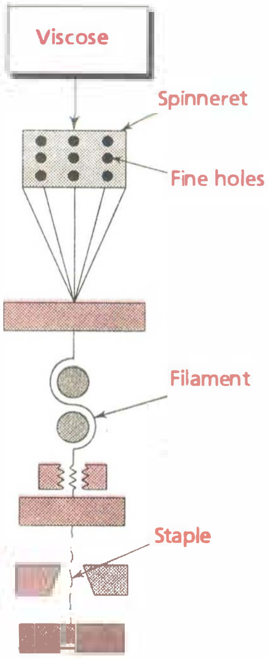


Fig. 8.1 Preparation of viscose rayon

Uses

- It is used to manufacture tyre cords.
- It is woven to make fabrics like sarees and mixed with wool to make carpets.
- It is used to make bandages and surgical dressings.
- It is mixed with cotton for making dresses, aprons and caps.

POLYESTER

Polyester was first prepared in England. It is a general term used for fibres which contain many ester groups. *Terylene*, *Dacron*, are some examples of polyester fibres.

Properties

- Polyester fibres are very strong and wrinkle resistant.
- They absorb less water and dry easily.
- They are not attacked by moths and ordinary chemicals.

Uses

- They are used in the manufacture of textiles.
- Terylene when mixed with cotton and wool forms terrycot and terrywool, respectively. These are used to make shirts, trousers, suits, etc.
- They are used to make conveyor belts.

The various types of fibres can be distinguished by the *burning test*. Table 8.1 shows the results obtained when a thread of silk, wool, rayon, cotton, terylene and nylon are burnt.

Table 8.1 Testing fibres by burning

Fibre	On burning
1. Silk and wool	Burn slowly with a smell of burning feathers or hair. No bead formation.
2. Cotton	Burns vigorously with a smell of burning paper. No bead formation.
3. Rayon	Burns readily with a smell of burning paper.
4. Nylon and cashmilon	Burn slowly with a sooty flame. Bead is formed.
5. Terylene	Burns slowly with bead formation.

Remember . . .

- Fibres are thread-like materials which are strong and flexible enough to be converted into clothes, ropes, nets, etc.
- Cotton, wool, silk, linen are examples of natural fibres.
- Nylon, rayon, polyester and acrylics are some common synthetic fibres.
- Nylon, polyester and acrylics are made from plastics. Rayon is made from cellulose.

PLASTICS

No materials are more versatile than plastics. Plastic means 'easy to mould'. A plastic is a substance that can be moulded when soft and then hardened to produce a durable article.

Table 8.2 Common plastics and their uses

Plastic	Uses
Polyvinyl chloride (PVC)	raincoats, records, garden hoses and electrical insulation.
Polyamide (Nylon)	shirts, jumpers, tights, socks.
Polyester (Terylene)	clothes, seat belts, yacht sails.
Polypropene	string, carpets.
Polyurethane	furniture padding.
Polyvinyl acetate	glue and paint.
Polytetrafluoroethene (Teflon)	non-stick surfaces.
Polymethyl methacrylate (Perspex)	safety glass.
Melamine	children's cups and plates and kitchen shelves.
Bakelite	electrical fittings.

Plastics are *polymers* made up of thousands of small molecular units called *monomers.*

The process of linking together small molecules to form large molecules called polymers is called polymerization.

$$\text{Monomer} \xrightarrow{\text{polymerization}} (\text{Monomers})_n \text{ or polymer}$$

Plastics are made by heating, often under pressure, gases or liquids obtained from crude oil in oil refineries.

Table 8.3 Advantages and disadvantages of plastics

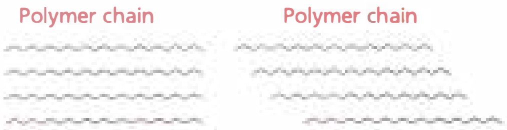
Advantages	Disadvantages
1. Quite cheap and easily made.	1. They are not easily disposable. Plastic bags and cartons do not rot when they are disposed off, so they pollute the atmosphere. But there are some plastics which are biodegradable.
2. Lighter than wood, stone or metal.	2. Some plastics are flammable ie, they catch fire easily.
3. Unreactive. They do not corrode in air or water. Many are not affected by acids or alkalis.	3. They produce harmful gases on burning. For example, PVC gives off fumes of hydrochloric acid on burning; carbon monoxide and hydrogen cyanide are released when polyurethane burns.
4. Do not conduct heat or electricity, so can be used as insulators.	4. They usually do not look as good as wood or stone.
5. Can be moulded into any shape.	
6. Can be made very strong.	
7. Can be coloured, by adding pigments.	

Groups of plastics

There are two groups of plastics, *thermoplastics* and *thermosetting plastics* or *thermosets.*

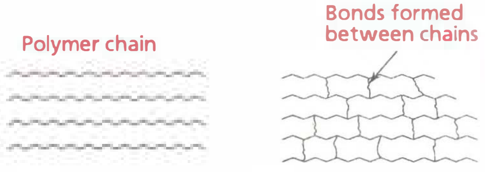
Thermoplastics These plastics get soft and runny when they are heated, and hard again when they are cooled. They can be made soft and hard over and over again, eg, polyethene, PVC, nylon, polyester, acrylic plastics.

Thermosets They do not soften on heating and can be moulded only once, eg, melamine, urea, bakelite (phenol), methanol.



(a) Polymer chains lie next to each other (b) The chains can slide past each other

Fig. 8.2 Thermoplastic



(a) Polymer chains lie next to each other (b) Heat causes bonds to form between chains

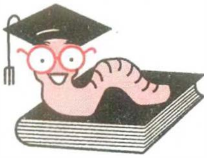
Fig. 8.3 Thermosets

Let's Answer . . .

1. Why are thermoplastics not used for making handles for cooking utensils?
2. Classify the following plastics as thermoplastics and thermosets:

   PVC, melamine, bakelite, polyester, polythene, urea, phenolics, methanol, acrylics, nylon

Remember . . .



- Plastics are synthetic materials that can be moulded when soft and then hardened to produce a durable article.
- Thermoplastics melt easily on heating and can be moulded over and over again on repeated heating, eg, PVC.
- Thermosets do not soften on heating and can be moulded only once, eg, melamine.

GLASS

Glass is an amorphous transparent substance obtained by the solidification of a liquid. It looks like a solid but is a super-cooled liquid. It is a mixture of a number of silicates. It does not have any definite composition.

Manufacture

Ordinary glass is a mixture of sodium silicate, calcium silicate and silica. The raw materials used in the manufacture of ordinary glass are sodium carbonate (Na_2CO_3), limestone ($CaCO_3$), silica (SiO_2).

Glass is prepared by heating a mixture of Na_2CO_3, $CaCO_3$ and silica in a furnace to a temperature of about 1400 °C. A clear and transparent jelly-like substance is formed. The product is subjected to slow cooling (annealing) to form glass, which is then moulded into different shapes.

Different chemicals are added to make different glasses, eg, calcium phosphate to make opalescent glass, cobalt oxide to make blue glass, boron oxide to make borosilicate glass, etc.

Types of glass

Some of the common types of glass are described below.

Soda glass This is the most common type of glass. It is prepared from sodium carbonate (soda ash) and so it is called *soda glass*.

Pyrex It is made by fusing silica, boric oxide and aluminium oxide. It is also called *borosilicate glass*. It is resistant to heat, chemicals and shock. It is used in laboratory and kitchen glassware.

Float glass It is made by floating molten glass on huge beds of molten tin so that it is absolutely flat when it hardens. It is used for making windows in shops.

Flint glass It is a clear glass made from potassium and lead silicate. It is soft and transparent and used for making lenses, prisms and other optical instruments.

Safety glass is made by placing a sheet of plastic between sheets of glass plates. It is used in automobiles and bullet-proof screens.

Lead glass is used for making shields against radioactivity and in expensive drinking glasses.

Photochromatic glass is a special type of glass that darkens when exposed to sun or bright light.

Glass wool is a fibre types of glass used for filtration and refrigeration. It is also used as an insulating material.

CERAMICS

The word *ceramic* is derived from the old Greek word *keramos*, meaning *potter's clay*. They are compounds of silicon, carbon, nitrogen and oxygen, made from clay, feldspar and flint.

Manufacture

Clay used in the manufacture of ceramics is chiefly hydrated aluminium silicate ($Al_2O_3.SiO_2\ 2H_2O$). These silicates obtained from the earth's crust are crushed, grounded and sieved to produce fine and better variety of clay. Water is added to the clay to make a uniform dough by kneading the mixture. This dough is then moulded into desired shapes of objects and dried. The dried product is heated in a special type of furnace (called kiln) to about 800 °C. Heating provides strength to the product. The product is either used as such or is covered with a coating of glaze to make it smooth and resistant towards liquid absorption. Coloured ceramics are obtained by adding coloured chemicals.

Properties

- They are resistant to the action of chemicals like acids, alkalis, oils, etc.
- They are bad conductors of heat and electricity
- They are stable at high temperatures.

Uses

- They are used to make cups, saucers, plates, bowls, and flower pots.
- They are used to store pickles and jams as they are resistant to chemicals.

- They are used in making electrical insulations for power lines, spark plugs.
- They are used for making roof tiles, bricks, etc.
- Glazed ceramic is used to make tiles, sinks, washbasins, etc. It forms the lining of furnaces as it is heat resistant.
- Silicon carbide (a ceramic) is used to coat knives and scissors.

Remember . . .

- Glass is an amorphous substance obtained by the solidification of a liquid. It is a super-cooled liquid.
- Sodium carbonate, limestone, and silica are the raw materials used in the manufacture of glass.
- Ceramics are compounds of silicon, carbon, nitrogen and oxygen, made from clay, feldspar and flint.

SOAPS

Soaps are sodium salts of fatty acids such as stearic acid, oleic acid and palmitic acid. They are used for cleaning clothes, body, utensils, etc.

Manufacture

The raw materials used in the manufacture of soap are

(i) animal fat or vegetable oil
(ii) caustic soda (NaOH)
(iii) sodium chloride (NaCl)

Animal fats such as beef tallow or vegetable oils such as palm oil and coconut oil are heated with caustic soda solution. The fat or oil react with caustic soda to form soap and glycerol.

$$\underset{\text{(ester)}}{\text{Fat or Oil}} + \underset{\text{(sodium hydroxide)}}{\text{Caustic soda}} \longrightarrow \underset{\text{(sodium salt of fatty acid)}}{\text{Soap}} + \underset{\text{(alcohol)}}{\text{Glycerol}}$$

The alkaline hydrolysis of fats and oils to make soaps is called saponification.

After saponification, sodium chloride is added to the solution to cause the precipitation of soap. Solid soap floats on the surface. The crust of soap is removed and moulded to different shapes. Various colouring agents and perfumes are added depending on the requirement.

Activity 1

Aim To prepare soap.

Materials required Beaker, evaporating dish, castor oil, concentrated NaOH solution, glass rod, distilled water, tap water, burner, salt solution.

Method Half fill a beaker with tap water and boil it. Take 2 cc of castor oil and 10 cc of concentrated NaOH solution in an evaporating dish and place it on top of the beaker. Stir the mixture for 10-15 minutes. Now add saturated salt solution to the dish and stir it. Remove the dish from the hot beaker and cool it for sometime. A layer of soap is formed on the surface (Fig. 8.4).

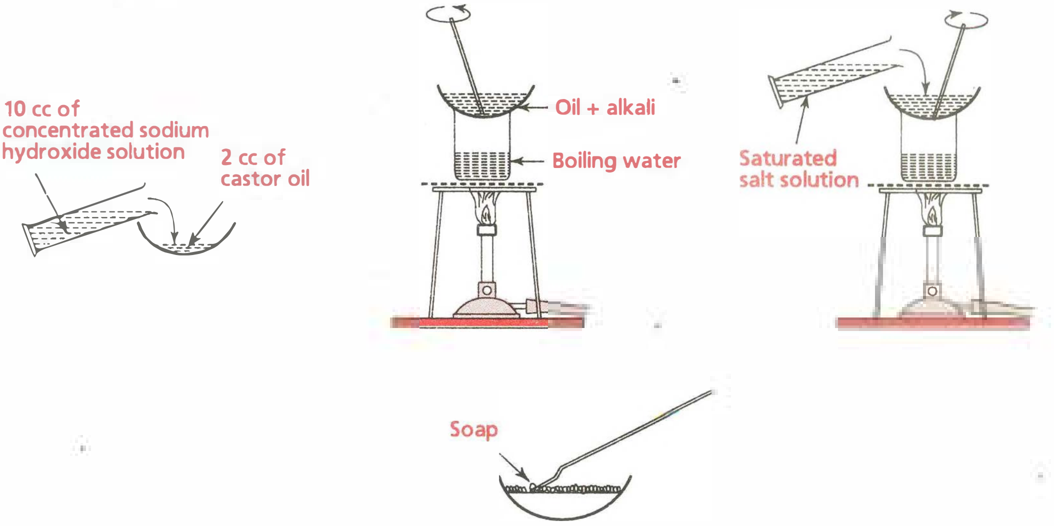


Fig. 8.4 Formation of soap from vegetable oil and caustic soda

Cleansing action of soap A soap molecule has two ends. One end of the soap molecule lowers the surface tension of water, thereby spreading the material being washed. The other end emulsifies oil and grease and dislodges the dirt particles (Fig. 8.5).

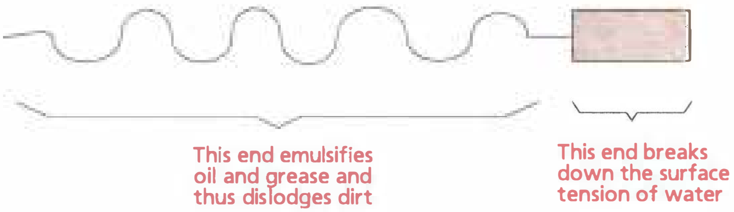

This end emulsifies oil and grease and thus dislodges dirt
This end breaks down the surface tension of water

Fig. 8.5 A soap molecule

Action of soap on hard water Ordinary soap is sodium stearate ($C_{17} H_{35}COONa$). When hard water is added to soap, the calcium and magnesium ions present in hard water react with sodium stearate to form insoluble precipitate of calcium and magnesium stearate. This is known as *scum*. The scum clings to the pores of clothes and chokes them, due to which the clothes are not cleaned properly.

$$\underset{\text{(soap)}}{\text{Sodium stearate}} + \underset{\text{(Ca and Mg ions)}}{\text{Hard water}} \longrightarrow \underset{\text{(scum)}}{\text{Calcium stearate} + \text{magnesium stearate}}$$

Activity 2

Aim To distinguish between soft and hard water.

Materials required Three test tubes, soap solution, tap water, river water, well water.

Method Take tap water, river water, and well water in three test tubes. Add soap solution to each and shake well. What do you observe? Soap produces lather with tap water. No lather is formed with well and river water.

This activity shows that tap water is *soft water*, since it forms lather with soaps readily. River and well water are *hard water*, since they do not form lather with soaps.

Advantages of soaps

- They are biodegradable (decomposed by micro-organisms).

Disadvantages of soaps

- They are not suitable for washing with hard water.
- They are less soluble in water.
- They have a weak cleaning action.
- They consume a large quantity of vegetable oil during their manufacture.

DETERGENTS

Detergents are synthetic chemicals obtained from petroleum hydrocarbons. The first synthetic detergent was made in France in 1831. Detergents are more soluble in water than soaps. They are called *soapless soaps* because they do not form insoluble scum with hard water and can be readily used with it.

Sodium lauryl sulphate ($C_{12}H_{25}OSO_3Na$) is an example of a synthetic detergent. It consists of two parts--hydrophobic part (water repellent) and hydrophilic part (water attracting).

Advantages of Detergents

- They can be used even with hard water.
- They are prepared from petroleum hydrocarbons and thus save vegetable oils for human consumption.
- They can act in acidic media also.
- They are better cleaning agents and more soluble in water than ordinary soaps.

Disadvantages of Detergents

- They are non-biodegradable.
- They cause water pollution.

The differences between soaps and detergents is summarized in table 8.5.

Table 8.5 Soaps and detergents—A comparison

	Soaps	Detergents
1.	They are sodium salts of fatty acids	They are sodium salts of petroleum hydrocarbons.
2.	They are prepared from animal fat or vegetable oils.	They are prepared from petroleum hydrocarbons.
3.	They are suitable for washing with hard water.	They can be used with hard water.
4.	They are biodegradable.	They are not biodegradable.
5.	They have a weak cleansing action.	They have a strong cleansing action.

Remember . . .

- Soaps are sodium salts of fatty acids.
- Animal fats and concentrated sodium hydroxide are used in the manufacture of soap.
- Detergents are synthetic chemicals made from petroleum hydrocarbons. They are called soapless soaps.

FERTILIZERS AND MANURE

Plants need certain essential elements for their growth. Most of these elements are acquired by the plants from air, water and soil (eg, C, H, O).

The most important elements required for plant growth are nitrogen, phosphorus and potassium, which are called *major* or *primary nutrients*. Other elements like iron, manganese, boron, copper, zinc, molybenum are need by the plants in small quantities. They are called *micronutrients* or *trace elements.*

Soil is rich in nutrients and plants absorb them from the soil. However, persistent cropping depletes the soil of its nutrient content. The deficiency of the necessary nutrients in soil is replenished by adding *manure* and *fertilizers.*

Manure is a natural substance obtained by the decomposition of animal wastes and plant residues.

They provide essential nutrients and humus to the soil and maintain its fertility. They are also called *organic fertilizers*.

Fertilizers are water soluble compounds rich in one or more of the essential nutrients.

They enrich the soil with nutrients and make it more fertile. They are also called *inorganic* or *artificial fertilizers.*

Fertilizers are classified under three categories depending on the type of major nutrient present in them. They are:

(i) Nitrogenous fertilizers.
(ii) Phosphorus fertilizers.
(iii) Potassium or Potash fertilizers.

Sometimes, certain chemicals are added to the soil to make the nutrients easily available to plants. Such chemicals are called *stimulant fertilizers*, eg, lime and gypsum.

Nitrogenous fertilizers Nitrogen is the most essential requirement of plants. However, only the leguminous plants like algae, fungi, mosses, etc, have the capacity to absorb the freely available atmospheric nitrogen. In order to make nitrogen easily available to plants, fertilizers rich in nitrogen are added to the soil. Some common nitrogenous fertilizers are: Ammonium sulphate $(NH_4)_2SO_4$; Ammonium nitrate (NH_4NO_3); Urea (NH_2CONH_2); Calcium ammonium nitrate $(Ca(NO_3)_2\ NH_4NO_3)$.

Phosphorus fertilizers Phosphorus is needed by the plants for cell division. It promotes root development and improves the quality of fibres. It hastens the ripening of fruits and improves the yield of food grains.

A few important phosphorus fertilizers are listed below:

(i) Ammonium phosphate : $(NH_4)_3\,PO_4$
(ii) Superphosphate : $Ca(H_2PO_4)_2\ 2CaSO_4$
(iii) Phosphatic slag : $Ca_3(PO_4)_2$ and $CaSiO_3$
(iv) Triple superphosphate : $3Ca(H_2PO_4)_2$
(v) Ammophos : $(NH_4)H_2PO_4$

Potash fertilizers Potassium is essential for plant growth. It enables the plant to have an effective root system. It helps in the formation of albuminoids and carbohydrates in plants. It makes the plant stronger and disease resistant. Some important potash fertilizers are:

(i) Potassium chloride : KCl
(ii) Potassium nitrate : KNO_3
(iii) Potassium sulphate : K_2SO_4
(iv) Potassium carbonate : K_2CO_3

Fertilizers containing only one primary nutrient (nitrogen, phosphorus or potassium) are called *single fertilizers*, eg, ammonium sulphate, while those containing more than one primary plant nutrient are called *mixed fertilizers*, eg, ammophos, potassium nitrate, etc. Fertilizers containing all the three primary nutrients (N, P and K) are called *complete fertilizers* or NPK *fertilizers*.

Some common fertilizer factories in India is shown in table 8.6.

Table 8.6 Fertilizer factories in India

Location	State
Sindri	Bihar
Nangal	Punjab
Trombay	Maharashtra
Gorakhpur	U.P.
Durgapur	West Bengal
Barauni	Bihar
Alwaye	Kerala

Precautions while using fertilizers

- Presence of excessive fertilizers in the soil make it poisonous. Hence a right quantity of fertilizer should be added.
- Proper irrigation of the soil is essential before the addition of fertilizers.
- Lime should not be added with ammonium fertilizers as lime traps nitrogen.
- Fertilizers like urea should be stored in a dry place.
- Nitrate fertilizers should be stored away from coal and sawdust to avoid fire.

PESTICIDES, FUNGICIDES AND HERBICIDES

Pesticides are chemicals added to control pests like rats, insects, etc, which damage the crops. They are sprayed in the form of dust or as suspensions in a liquid. Some common pesticides are:

- DDT (Dichlorodiphenyl-trichloroethane)
- BHC (Benzene hexachloride)
- Methyl parathion.
- Heptachlor.
- Dieldrin
- Chlordane
- Malathion.

Requirements of a good pesticide

- It should be biodegradable.
- It should be cheap and easy to use.
- It should be selective in action. It should be effective against harmful pests but not harmful to human beings, useful birds, bees and animals.

Precautions while using pesticides

- It should be kept out of the reach of children.
- Eyes, nose and mouth should be protected, while spraying them.
- It should not be used in excess.
- It should be kept away from food articles.
- Domestic animals and birds should not be allowed in the field during pesticide treatment.

The chemicals added to kill fungi which cause plant diseases are called fungicides.

Some of the important fungicides are:

- Bordeaux mixture (Mixture of $CuSO_4$ and CaO)
- Formalin (Aqueous solution of formaldehyde)
- Powdered sulphur.
- Sulphur dioxide (SO_2)

The chemicals used to destroy weeds (unwanted plants) are called herbicides. They can be harmful to the useful plants also, so one has to be careful while using these chemicals. Some of the common herbicides are:

- I.P.C (Isopropyl *n*-phenyl carbamate)
- 2, 4, –D (2, 4-dichloro phenoxyacetic acid)
- 2, 4, 5, –T (2, 4, 5-trichloro phenoxyacetic acid)

Remember . . .

- Manure is a natural substance obtained by the decomposition of animal wastes and plant residues.
- Fertilizers are water soluble compounds rich in one or more of the essential nutrients. They are classified into three types—nitrogenous, phosphorus and potash fertilizers.
- Pesticides are chemicals added to kill pests which damage the crops, eg, DDT.
- Fungicides are chemicals used to kill fungi, eg, Bordeaux mixture.
- Herbicides are chemicals used to kill unwanted plants (weeds), eg, 2, 4 –D.

Questions

Answer in brief

1. What are man-made materials? Give examples
2. How are fibres classified? Give examples in each.
3. Explain the process involved in the manufacture of (a) viscose rayon (b) glass (c) soap.
4. Give the properties and uses of any two synthetic fibres.
5. Distinguish between soaps and detergents.
6. Explain cleansing action of soap. How does soap react with hard water?
7. Compare the properties of nylon, rayon, wool and cotton on the basis of burning test.
8. Write a brief note on pesticides and fungicides.
9. What are fertilizers? How are they classified? Explain.
10. What are the different types of glass?

Fill in the blanks

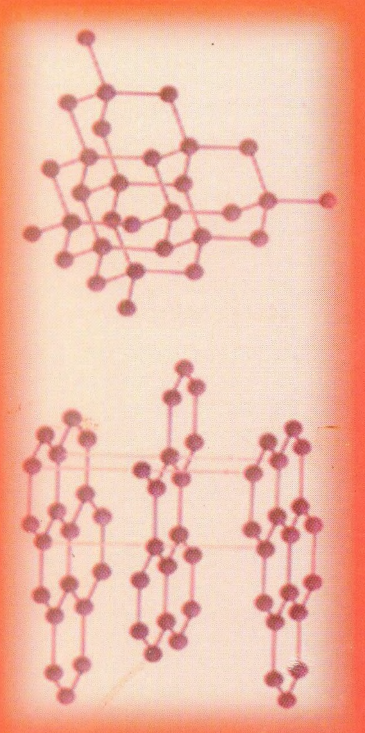
1. Melamine is a _____ plastic.
2. _____, _____ and _____ are used for making ceramics.
3. _____ is a complete fertilizer.
4. Rayon is also known as _____.
5. Nylon was discovered by _____.
6. _____ glass is used in automobiles.
7. Alkaline hydrolysis of fats and oils is called _____.
8. Cotton, silk, linen and wool are _____ fibres.
9. Rayon is made from _____.
10. 2,4 –D is an example of a _____.

Choose the correct alternative

1. Which of the following is a thermoplastic?
   (a) melamine (b) phenol (c) PVC (d) bakelite
2. Bordeaux mixture is a
   (a) pesticide (b) fungicide (c) herbicide (d) fertilizer
3. Which does not form lather with soaps?
   (a) well water (b) river water (c) tap water (d) rain water
4. Which is obtained from cellulose?
   (a) nylon (b) polyester (c) acrylics (d) rayon.
5. Which is not a major nutrient?
   (a) nitrogen (b) zinc (c) phosphorus (d) potassium.

Think & Answer . . .

1. Fertilizers can cause water pollution. How?
2. Why does soap produce scum in hard water?
3. Which plastic material would you choose for making these articles:
    - (a) Outdoor water pipes?
    - (b) Gramophone records?
    - (c) Windows?
    - (d) Cinema film?
    - (e) Switches?



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ISBN 0-19-564353-4

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